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ABSTRACT

The first part of this curriculum guide for a college-level general chemistry course includes: (1) a list of 28 lectures/lessons with topic titles and content divisions; (2) behavioral objectives related to specific lessons; (3) a list of laboratory activities and objectives; (4) a course overview and syllabus for spring semester 1981; and (5) a class schedule organized by week, lecture number, and lecture title. The major portion of the guide is devoted to individual lessons, including for each lesson: lesson number; title; instructional time; behavioral objectives (with pertinent conditions and performance criteria); student and instructor references and materials/equipment; detailed lecture outline; student activities; notes on application of lesson material to other areas, summary, evaluation, and homework; and handouts and student assignments such as an introduction to the periodic table following Lesson 5. General topics related to lessons include: fundamental concepts of chemistry (Lesson 1); structure of the atom and properties of matter (Lessons 2-5); chemical bonding (Lessons 6-7); conservation of mass and related problems (Lessons 11-14); balancing of chemical equations (Lessons 15-19); chemical states (Lesson 23); and acids/bases (Lesson 26). Other lessons provide information for examination reviews. Copies of examinations are also included. (JN)

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INSTRUCTOR'S GUIDE FOR GENERAL CHEMISTRY

by

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SE 035 804

FORWARD

General Chemistry is a three semester hour (one semester) post secondary course consisting of two one-hour lectures and one two-hour laboratory session each week. Assumptions are not made about entering students; that is, students taking general chemistry need not have taken high school chemistry prior to enrolling in this course. The course, however, does include a test out program. Students with extensive high school backgorunds in high school chemistry may, at their option, elect to take the test out examination. Additional test out information is provided in the test out program section of this guide.

The majority of lectures include an activity session. Those sessions are intended to be used by students to test their understanding of material presented during a lecture. Whenever possible, the activities were designed to be as concrete as possible and Piagetian based.

This guide is intended to be an instructor's tool. It includes all information necessary to teach the course as it was field tested.

COURSE SCOPE

The following subjects are covered during the course:

LECTURE/ SESSION NUMBER	TOPIC TITLE AND OUTLINE
1	Some fundamental concepts of chemistry A. Classification of matter 1. elements, compounds, mixtures 2. chemical and physical properties 3. chemical and physical changes B. Matter and energy 1. specific gravity, specific heat, density C. Measurements in chemistry
2	Structure of the atom and the periodic table A. Electrons, protons, neutrons B. Atomic numbers
3	Periodic law and the structure of the atom A. Isotopes B. Electronic structure and the periodic law C. Subshells, orbits and quantum numbers
4	Periodic law and the structure of the atom II A. Suborbital filling order B. Relationship between suborbital filling order and energy levels C. Valence electrons D. The octet rule
5	The relationship of periodic classification to the properties of the elements A. Variation of properties within the periods and groups of the periodic table B. Uses of the periodic table
6	General concepts used in chemical bonding A. Oxidation numbers

LECTURE/ SESSION NUMBER	TOPIC TITLE AND OUTLINE
	B. The molecule (ionic and covalent bonds)
	C. Electronegativities of elements
7	General concepts used in chemical bonding II
	A. Use of quantum numbers in chemical bonding theory
	B. Sigma and pi bonds
8	Review for Examination One
9	Examination One
10	Conservation of Mass in Chemical Systems
	A. Chemical formulae and equations
	B. The Avagadro number
	C. Gram molecular weights
	D. Percent composition from molecular formulae
11	Conservation of mass in chemical systems II
	A. The mole and balanced chemical equations
12	Conservation of Mass in Chemical Systems III
	A. Molarity
	B. Normality
	C. Equivalent weights
	D. Mole fraction
	E. Molarity
	F. Weight percent
	G. Limiting reagents
13	Problems and the Conservation of Mass
	A. Reactions by mass
	B. Limiting reagent problems
	C. Atoms to mass problems
	D. Atoms of A to grams of B problems

LECTURE/ SESSION NUMBER	TOPIC TITLE AND OUTLINE
14	Problems and the Conservation of Mass II
	A. Quantitative information form balanced equations
15	Balancing Chemical Equations I
	A. Electron gain and loss
	B. Oxidation and Reduction
	C. Balancing by inspection
16	Review for Examination Two
17	Examination Two
18	Balancing Chemical Equations II
	A. Oxidation-reduction equations 1. oxidizing agents 2. reducing agents 3. electronegativity and oxidation-reducation 4. balancing equations by the oxidation number method
19	Balancing Chemical Equations III
	A. Ion Electron Method (half reactions)
	B. Algebraic Method
20	Chemical States
	A. Solids B. Liquids C. Gases
21	The Gaseous State and Kinetic Molecular Theory
	A. The Avagadro Hypothesis B. Dalton's law C. Doyle's law D. Charle's law E. The Ideal gas law
22	Interrelationships of the Gas Laws
23	Solutions
	A. Definitions related to solutions B. Concentrations of solutions C. Mixing solutions D. Dilution of solutions E. Reactions and solution strength

LECTURE/ SESSION NUMBER	TOPIC TITLE AND OUTLINE
24	Review for Examination Three
25	Examination Three
26	Acids and Bases A. Properties of Acids and Bases B. Acid-base Neutralization C. pH and pOH
27	Final Examination

COURSE BEHAVIORAL OBJECTIVES

The course behavioral objectives are a list of competencies students should be able to demonstrate as a result of having taken the course. They are the areas around which evaluations of student competencies revolve and further, around which any testing of students is designed. Conditions under which minimum competency with each objective must be demonstrated are listed in each lesson plan. The criteria reflecting minimum competency levels are also included on the individual lesson plans.

The objectives are listed by lesson number and then by sequence of coverage within each lesson.

LESSON NUMBER	OBJECTIVE NUMBER	OBJECTIVE
3	I	Define density, specific gravity and specific heat
	II	Make measurements in chemistry
	III	Describe chemical and physical changes in matter
	IV	Identify elements compounds and mixtures
	V	Differentiate between chemical and physical properties of matter
4	I	Describe the structures of atoms in terms of electrons, protons and neutrons
	II	Draw theoretical pictures of atoms based upon atomic numbers
5	I	Define isotopes
	II	Write orbital formulae for the first 54 elements
	III	Draw s and p subshells
	IV	Relate suborbitals to the structure of the atom
6	I	Predict the suborbital filling order for the majority of elements on the periodic table
	II	Relate suborbital filling order to energy levels
	III	Define valence electrons
	IV	Define the octet rule and its relationship to molecular bonding theory
7	I	Describe ways in which elements change as one proceeds from right to left across a period
	II	Describe differences between elements on the left and right sides of the periodic table
	III	Describe ways in which elements change as one proceeds from top to bottom of a group or family
	IV	Describe uses of the periodic table
8	I	Predict molecular formulae using common oxidation numbers
	II	Describe the difference between ionic and covalent character of bonds

LESSON NUMBER	OBJECTIVE NUMBER	OBJECTIVE
8	III	Determine the degree of ionic character based upon electronegativities
9	I	Identify principle and ℓ quantum numbers of electrons involved in a bonding
	II	Relate bonding electrons to bond types
	III	Draw sigma and pi bonds
10	I	Identify misconceptions with course material covered since the beginning of the course
11	I	Identify areas of conceptual weakness with course material covered heretofore
12	I	Define molecular, empirical or simplest and structural formulae
	II	Define chemical equation
	III	Relate the Avogadro number to gram molecular weights and the periodic table
	IV	Compute molecular mass and composition from molecular formulae
13	I	Describe the meaning of conservation of mass in terms of (1) molecular coefficients; (2) moles of molecules; (3) mass in grams
14	I	Relate the following concepts to chemical equations: (1) molarity; (2) normality; (3) equivalent weights; (4) molarity; (5) weight percent; (6) limiting reagent
15	I	Solve reactions by mass problems
	II	Solve atoms to mass problems
	III	Solve mass to moles problems
	IV	Solve atoms A to grams of B problems
16	I	Solve quantitative problems based upon balanced chemical equations
17	I	Relate electron gain or loss to oxidation or reduction
	II	Define oxidation and reduction
	III	Balance simple chemical equations by inspection

LESSON NUMBER	OBJECTIVE NUMBER	OBJECTIVE
18	I	Identify misconceptions with course materials covered between examinations 1 and 2
19	I	Identify conceptual weaknesses with course material covered between examinations 1 and 2
20	I	Identify oxidizing and reducing agents in chemical reactions
	II	Describe the relationship between electronegativity and oxidation-reduction
	III	Balance chemical equations by the oxidation number method
21	I	Balance chemical equations by the ion electron method
	II	Balance chemical equations by the algebraic method
22	I	Describe time-temperature curves
	II	Compute heats of fusion and vaporization based upon empirical data
23	I	Describe the Avagadro hypothesis
	II	Solve Dalton's, Boyle's Charles' and Ideal Gas Law problems
24	I	Solve Gas Law problems
25	I	Define terms associated with solutions
	II	Describe how solutions of specific strengths are mixed
	III	Describe mathematically how to dilute solutions of known strength
	IV	Relate solution strength to chemical reactions
26	I	Identify misconceptions with course material covered between examinations 2 and 3
27	I	Identify areas of conceptual weaknesses with course material covered between examinations 2 & 3
28	I	Define acids,bases and terms relatesd to acids and bases

LESSON NUMBER	OBJECTIVE NUMBER	OBJECTIVE
28	II	Describe the functions of indicators in acid-base neutralizations
	III	Solve acid-base neutralization problems based upon balanced chemical equations
29-30	I	Identify course-wide misconception regarding basic chemistry

LABORATORY OBJECTIVES AND ACTIVITIES

The laboratory and lecture portions of General Chemistry (CH-1) are an integrated whole. The following list of student centered behavioral objectives have been established for the lab portion of the course. Students should attain these objectives as a result of completing the laboratory activities which follow the objectives. Student facilities with the objectives are evaluated by examining the laboratory reports which accompany each lab activity. Students who answer correctly 90% of all lab report questions will have demonstrated successful attainment of all the objectives.

OBJECTIVES

As a result of successfully completing all of the following laboratory activities, students should be able to:

1. Describe the qualitative and quantitive aspects of common laboratory equipment;
2. Outline the procedures used in evaluating physical properties of matter;
3. Use the physical properties of matter to identify specific substances;
4. Use decantation, extraction and sublimation techniques to separate substances from one another;
5. Observe chemical reactions and identify some products of those reactions;
6. Summarize chemical changes in terms of balanced chemical equations;
7. Determine emperical formulae;
8. Discuss the properties of a typical transition element;
9. Determine percent yeild from a chemical reaction;
10. Identify some common substances found in everyday life;

11. Determine the relativities of metals in chemical reactions;
12. Describe the relationship between temperature and pressure in a fixed volume system;
13. Determine the values for the faraday and Avagadro's number by electrolysis;
14. Describe colligative properties;
15. Use Colligative properties of a substance to determine the molecular weight of a substance;
16. Use titration to determine the amount of acid in an unknown;
17. Make observations and interpret and draw conclusions from observed facts;
18. Keep satisfactory records and write satisfactory lab reports;
19. Draw conclusions only from observable facts or accepted data.

LABORATORY ACTIVITIES

Students enrolled in General Chemistry (Ch-1) conduct the following laboratory activities:

1. Basic laboratory techniques;
2. Identification of substances of physical properties;
3. Separation of components of a mixture;
4. Chemical reactions;
5. Chemical formulas;
6. Chemical reactions of copper and percent yield;
7. Chemicals in everyday life: what are they and how do we know?
8. Behavior of gases; molecular weight of a vapor;
9. Activity series;
10. Electrolysis, the faraday and Avagadro's number;
11. Colligative properties; freesing point depression and molecular weight;
12. Titration of acids and bases.

Ch-1 GENERAL CHEMISTRY

Course Overview — Spring Semester 1981

Materials

A. Required
   1. Brown, T.L. and LeMay, F. Chemistry: The Central Science (2nd edition). Englewood Cliffs, New Jersey: Prentice-Hall Inc. 1980.

   2. Nelson, J.H. and Kemp, K.C. Laboratory Experiments for Brown and Lemay Chemistry the Central Science. Englewood Cliffs, New Jersey: Prentice-Hall, Inc. 1977.

B. Recommended-Optional

   1. Hill, J.C. Students Guide to Brown and LeMay Chemistry: the Central Science. Englewood Cliffs, New Jersey: Prentice-Hall Inc. 1980.

Laboratory Schedule

SEMESTER WEEK	DATE WEEK BEGINS	LABORATORY NUMBER
1	Jan. 26	no labs
2	Feb. 02	1
3	Feb. 09	2
4	Feb. 16	3
5	Feb. 23	4
6	Mar. 02	5
7	Mar. 09	6
8	Mar. 16	7
9	Mar. 23	13
10	Mar. 30	16
11	Apr. 06	11 & 14
12	Apr. 13	17
13	Apr. 20	Lab Exam
14	Apr. 27	Final exams

Course Grades

Lecture Exams (3)	100 points each	300
Final Exam		200
Lab Exam		120
Labs (12)	15 points each	180
Total		800

Final Grades	A-720-800 B-640-719 C-560-639 D-480-559

Examination Dates

1. Lecture examinations will be given during the Tuesday class periods of weeks 5, 9 and 13.

2. The laboratory examination will be given during the regular scheduled laboratory periods of the last week of classes.

3. The final examination will be given during the final exam week.

Course Policy

1. Attendance at laboratory sessions is mandatory. Credit will not be given to those failing to attend. Specials do not relieve you of your responsibility to attend labs and since lab sections are full, lab makeups during other than your normally scheduled session will not be possible.

2. If you have to be on watch when an examination is given, you must notify your instructor 24 hours in advance of the examination if you desire to take a makeup exam. When this is done you must present your instructor with a copy of the Academy watch bill listing you as a watchstander. Makeup exams will under no conditions be given to those who fail to follow these procedures.

3. Makeup exams will not be given to people on special during an exam period.

4. The use of profanity in and around chemistry lectures and labs is not acceptable. Such practice will lead to dismissal from Ch1.

5. Students who have previously taken introductory college chemistry need not take this course. See your instructor for details.

6. Students who have taken two years of high school chemistry with final grades of C or better both years and those who took high school chemistry for one year and received a grade of A may test out of Ch1. See your instructor about test out details.

MAINE MARITIME ACADEMY
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DEPARTMENT OF ARTS AND SCIENCES

COURSE OUTLINE

CH 1 GENERAL CHEMISTRY

TIME FRAME: This course is presented to midshipmen for fourteen weeks in the second semester of the third class year.

TIME DISTRIBUTION: 2 class hours per week; 2 laboratory hours per week

CREDIT: 3 semester hours

TEXTBOOK: Brown, T.L., and LeMay, H.E. CHEMISTRY: THE CENTRAL SCIENCE (2nd Ed.). Englewood Cliffs, New Jersey: Prentice Hall, Inc., 1980.

COURSE OBJECTIVES: As a result of taking this course, STUDENTS should be able to manifest an understanding of the basic principles of chemistry. These principles include, but are not limited to atomic structure; periodic properties of the elements; chemical bonding; formulas and equations; properties of gases, solids and liquids; solutions, oxidation-reduction and acids and bases.

FIRST WEEK: **Lecture 1.** Some fundamental concepts of chemistry.

A. Classification of matter
   1. elements, compounds and mixtures
   2. chemical and physical properties
   3. chemical and physical changes

B. Matter and energy
   1. density, specific gravity and specific heat

C. Measurements in chemistry

Assignment:

A. Required Reading - Brown & LeMay
   1. Chapter 2, Sections 2.1, 2.2, 2.3, 2.4
   2. Chapter 1, Sections 1.3, 1.5
   3. Chapter 4, Section 4.6

B. Problems: 1.35, 1.38, 1.40, 1.48, 2.4, 2.44a, b, d

C. Suggested Reading - Hill
   1. Chapter 1.

Lecture 2. Structure of the atom and the periodic table.

A. Electrons, protons, neutrons

B. Atomic numbers

Assignment:

A. Required Reading - Brown & LeMay
  1. Chapter 2, Sections 2.6, 2.7
  2. Chapter 20, Introduction, pp. 592-3.

B. Problems: Chapter 2, p. 56. Learning goals, answer items 6, 7, 9; 2.26, 2.48, 2.49

C. Suggested Reading - Hill
  1. Chapter 2

SECOND WEEK

Lecture 1. Periodic law and the structure of the atom.

A. Isotopes
B. Electronic structure and the periodic law
C. Subshells, orbits and quantum numbers

Assignment:

A. Required Reading - Brown & LeMay
  1. p. 43, 592
  2. Chapter 2, Section 2.6
  3. Chapter 6, Sections 6.1, 6.2, 6.3, 6.4
  4. Chapter 5, p. 136-141 quantum numbers and orbitals.

B. Problems: 2.24, 2.28, 6.11, 6.14, 6.40, 6.7.

C. Suggested Reading - Hill
Chapter 6.

Lecture 2. Periodic Law and the Structure of the Atom II

A. Electronic structure and the periodic law

B. Subshells, orbits and quantum numbers.

Assignment:

A. Required Reading - Brown & LeMay

  1. Chapter 2, Sections 2.7, 2.8
  2. Chapter 7, Sections 7.1, 7.7

B. Problems: 5.30, 5.34, 5.50, 5.51, 5.52.

Lecture 1. The relationship of periodic classification to the properties of the elements.

A. Variation of properties within the periods and groups of the periodic

B. Uses of the periodic table

Assignment:

A. Required Reading - Brown & LeMay
   1. p. 45, 160, 133 (ionization energy)
   2. Chapter 6, Section 6.8
   3. Chapter 21, section 21.1

B. Problems: 5.15, 5.32, 5.35, 5.57, 21.1, 21.14

C. Suggested Reading - Hill
   Chapters 5 & 21

Lecture 2. General Concepts of Chemical Bonding

A. Oxidation Numbers

B. The molecule (ionic and covalent bonds).

C. Electronegivity of elements.

Assignment:

A. Required reading - Brown & LeMay
   1. Chapter 2, Section 2.8.
   2. Chapter 7, Section 7.1, 7.2, 7.4, 7.5, 7.9, 7.10.

B. Problems: 2.28, 2.32, 2.52a, 2.57, 7.1, 7.2, 7.4, 7.7, 7.10

C. Suggested Reading - Hill

   Chapter 7.

FOURTH WEEK

Lecture 1. General Concepts of Chemical Bonding II.

A. Use of quantum numbers in chemical bonding.

B. Sigma and pi bonds.

Assignment:

A. Required Reading - Brown & LeMay

1. Chapter 8, Sections 8.3, 8.4, 8.5, 8.6.

B. Problems: 8.19, 8.38.

C. Suggested Reading - Hill,

1. Chapter 8.

Lecture 2. Review

Assignment:

A. Required Reading - Brown & LeMay

1. Review reading required to date.

B. Review problems solved to date.

C. Review dimensional analysis.

FIFTH WEEK

Lecture 1: Examination 1.

Lecture 2. Conservation of mass in chemical systems

A. Chemical formulae and equations

B. The avogadro number

C. Gram molecular weights

D. Percent composition from formulae.

Assignment:

A. Required reading - Brown & LeMay

1. Chapter 2, Section 2.8

2. Chapter 3, Section 3.4, 3.6, 3.7, 3.8, 3.9, 3.1, 3.2, 3.3.

B. Problems: 3.10, 3.12, 3.15, 3.22, 3.47.

C. Suggested Reading - Hill, Chapter 3.

SIXTH WEEK

Lecture 1: Conservation of mass in chemical systems II.

A. The mole and balanced chemical equations.

Assignment:

A. Problems: 3.2, 3.3, 3.7a, c, e; 3.4

Lecture 2: Conservation of mass in chemical systems III.

A. Equivalent weights

B. Mole fraction

C. Molality

D. Weight percent

E. Limiting Reagents

F. Normality

G. Molarity

Assignment:

A. Required Reading - Brown & LeMay

1. Chapter 12, Sections 12.1
2. Chapter 3, p. 82 (limiting reagent).

B. Problems: 12.1, 12.3, 3.11, 3.77.

C. Suggested Reading - Hill, Chapter 12.

SEVENTH WEEK

Lecture 1. Problems and the conservation of mass.

A. Reactions by mass

B. Limiting reagent problems

C. Atoms to mass problems

D. Atoms of A to grams of B problems

Assignment:

A. Required reading - Brown & LeMay

1. Sample exercise 3.5, p 73
2. Sample exercise 3.6, p 77
3. Sample exercise 3.7, p 77
4. Sample exercise 3.8, p 77
5. Sample exercise 3.9, p 78
6. Sample exercise 3.10, p 78
7. Sample exercise 3.11, p 78

B. Problems: 3.21, 3.28, 3.30

C. Suggested Reading - Hill
Chapter 3

Lecture 2. Problems and the conservation of mass II.

A. Quantitative information from balanced equations

Assignment

A. Required reading - Brown & LeMay

1. Sample exercise 3.13, p 82
2. Sample exercise 3.24, p 83
3. Sample exercise 3.15, p 84

B. Problems: 3.45, 3.69

C. Suggested reading - Hill
Chapter 3

EIGHTH WEEK

Lecture 1. Balancing chemical equations I

A. Electron gain and loss

B. Oxidation and reduction

C. Balancing by inspection

Assignment:

A. Required reading - Brown & LeMay

1. Chapter 3, Section 3.2
2. Sample exercise 3.1
3. P 554
4. Chapter 7, Section 7.10

B. Problems: 3.2, 3.3, 3.4, 3.7, 3.56, 7.36, 7.37

C. Suggested Reading - Hill, Chapter 3

Lecture 2. Review

Assignment:

A. Required reading - Brown & LeMay

1. Review required reading between exams 1 and 2

B. Review text book problems solved between exams 1 and 2.

NINTH WEEK

Lecture 1. Examination 2

Lecture 2. Balancing chemical equations II

A. Oxidation - reduction equations

1. oxidizing agents
2. reducing agents
3. electronegativity and oxidation - reduction
4. balancing equations by the oxidation number method

Assignment:

A. Required reading

1. Brown & LeMay
    1. p 201, review section
    2. p 197, sample exercise 7.11
    3. p 197, sample exercise 7.12
2. Kolb, D. the chemical equation part II: oxidation - reduction reactions. Journal of Chemical Education, 326, -31, 55 (s), 1978.

B. Problems: 7.38, 7.39

TENTH WEEK

Lecture 1. Balancing Chemical Equations III

A. Ion Electron Method (half reactions)

B. Algebraic Method

Assignment:

A. Required Reading

1. Brown & LeMay

    1. Chapter 19, Section 19.1

    2. p 557, sample exercise 19.1

2. Kolb, D. the chemical equation part II: oxidation - reduction reactions. *Journal of Chemical Education,* 326, 331, 55(5), 1978.

B. Problems: 19.1, 19.3, 19.4

C. Suggested reading - Hill Chapter 19.

Lecture 2. Chemical States

A. Solids

B. Liquids

C. Gases

Assignment:

A. Required reading - Brown & LeMay

1. Chapter 9, section 9.1
2. Chapter 4, section 4.3
3. Chapter 18, sections 18.2, 18.3
4. p 538, figure 18.5
5. p 335, figure 11.29
6. Chapter 11, sections 11.1, 11.2, 11.3, 11.8

B. Problems: 11.10, 11.12, 11.18, 11.20, 11.56, 11.58, 11.77

C. Suggest reading - Hill Chapter 11.

ELEVENTH WEEK

Lecture 1. The Gaseous State and Kinetic Molecular Theory

A. The avogadro hypothesis

B. Dalton's Law

C. Boyles Law

D. Charles Law

E. The Ideal Gas Law

Assignment:

A. Required reading - Brown & LeMay

 1. Chapter 9, Section 9.2, 9.3, 9.4, 9.5

B. Problems: 9.3, 9.8, 9.13, 9.14, 9.18

C. Suggested reading - Hill, Chapter 9

Lecture 2: Interrelationships of the Gas Laws.

Assignment:

A. Required reading - Brown & Le May

 1. Review the reading done for the previous lecture

B. Problems: 9.70, 9.21, 9.23, 9.27, 9.29, 9.15

C. Suggested reading - review Hill, Chapter 9

TWELTH WEEK

Lecture 1. Solutions

A. Definitions related to solutions

B. Concentrations of solutions

C. Mixing solutions

D. Dilution of solutions

E. Reactions and solution strength

Assignment:

A. Required reading - Brown & Le May

1. Chapter 3, section 3.9

2. Chapter 12, sections 12.1, 12.2, 12.3, 12.4, 12.7

B. Problems: 12.1, 12.3, 12.6

C. Suggested reading - Hill, Chapter 12

Lecture 2. Review

Assignment:

A. Required reading - Brown & LeMay

1. Review required reading between exams 2 and 3

B. Review text book problems solved between exams 2 and 3

THIRTEENTH WEEK

Lecture 1. Examination 3

Lecture 2. Acids and Bases

A. Properties of acids and bases

B. Acid-base neutralization

C. pH and pOH

Assignment:

A. Required reading - Brown & LeMay

1. pp 67-69

2. Chapter 15, section 15.1, 15.2, 15.3, 15.8

B. Kolb, D. Aids and Bases.

Journal of Chemical Education,

459, 464, 55(7), 1978.

C. Kolb, D. the pH concept.

Journal of Chemical Education,

49-53, 56(1), 1979.

FOURTEENTH WEEK

Final Examination

CLASS SCHEDULE

The course is broken down here by weeks, lecture numbers and lecture titles

WEEK #	LECTURE #	LECTURE TITLE
1	1	Course Overview
1	2	Course Pretest
2	3	Some Fundamental Concepts of Chemistry
2	4	Structure of the Atom and the Periodic Table
3	5	Periodic Law and the Structure of the Atom I
3	6	Periodic Law and the Structure of the Atom II
4	7	The Relationship of Periodic Classification to the Properties of the Elements
4	8	General Concepts of Chemical Bonding I
5	9	General Concepts of Chemical Bonding II
5	10	Examination 1 Review
6	11	Examination 1
6	12	Conservation of Mass in Chemical Systems I
7	13	Conservation of Mass in Chemical Systems II
7	14	Conservation of Mass in Chemical Systems III
8	15	Problems and the Conservation of Mass I
8	16	Problems and the Conservation of Mass II
9	17	Balancing Chemical Equations I
9	18	Examination 2 Review
10	19	Examination 2
10	20	Balancing Chemical Equations II
11	21	Balancing Chemical Equations III
11	22	Chemical States
12	23	The Gaseous State and Kinetic Molecular Theory
12	24	Interrelationships of the Gas Laws

CLASS SCHEDULE (continued)

WEEK #	LECTURE #	LECTURE TITLE
13	25	Solutions
13	26	Examination 3 Reveiw
14	27	Examination 3
14	28	Acids and Bases

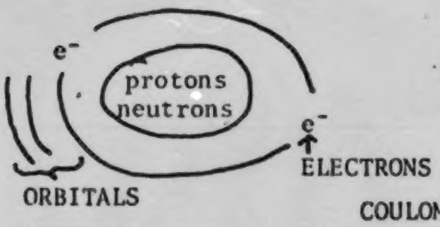
TITLE	Some Fundamental Concepts of Chemistry
LESSON NUMBER	1
TIME	50 minutes
DATE PREPARED	January 26, 1981
BEHAVIORAL OBJECTIVES	As a result of attendance at this lecture, students should be able to:
	I Define density, specific gravity and specific heat
	II Make measurements in chemistry
	III Describe chemical and physical changes in matter
	IV Identify elements, compounds and mixutres
	V Differentiate between chemical and physical properties of matter
I CONDITIONS	A. Using class notes and textbooks
	B. By the end of the class
II CRITERIA	Performance demonstrating minimum competency with the objectives is:
	A. Set up and or answer all of the questions correctly (those given by the instructor).
REFERENCES	
I INSTRUCTOR	A. Brown and Lemay (1.3; 1.5; 2.1; 2.2; 2.4; 4.6)
	B. Students guide to Brown and Lemay
	C. Solutions to Exercises in Brown and Lemay
II STUDENTS	A. "A" above
MATERIALS/ EQUIPMENT	
I INSTRUCTOR	A. Chalk
	B. Chalkborad
	C. Background sheets 1-4
II STUDENTS	A. Paper

PROCEDURE I. Lecture

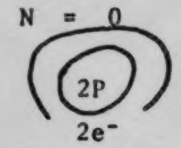
A. Atomic structure

1. Electrons, protons, neutrons, ions

a.


e⁻
protons
neutrons
e⁻
ELECTRONS
ORBITALS

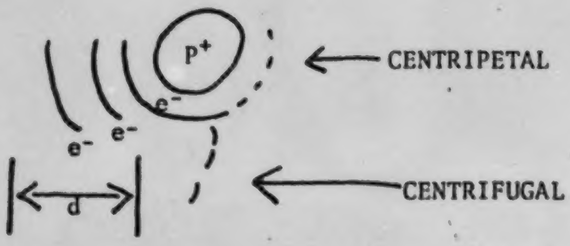
		COULOMBS
b. Charges	$e^- = -1$	-1.6×10^{-19}
	D = +1	$+1.6 \times 10^{-19}$
	N = 0	0



ELECTRICALLY NEUTRAL

c. All elements of the type have the same number of protons.

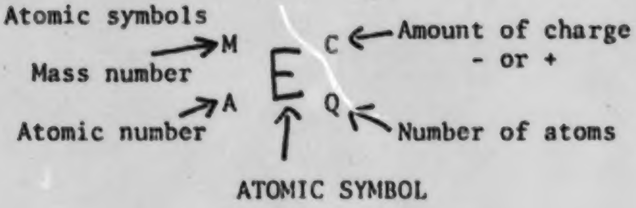
d.


P⁺
e⁻
e⁻
e⁻
d
CENTRIPETAL
CENTRIFUGAL

Attractive force $F = K \dfrac{Q_1 \; Q_2}{d^2}$

$K = 9 \times 10^9 \; Nm^2/C^2$

e. Atomic symbols


M
Mass number
A
Atomic number
E
C
Amount of charge - or +
Q
Number of atoms
ATOMIC SYMBOL

(1) at NO_o - number of protons

(2) mass NO_o - number of protons and neutrons or number of neucleons

(3) charge - exists only when an atom gains or looses an e^- (becomes ionized)

(4) Q exists whenever there are two or more atoms of the same element present

(5) Atomic symbol - identifies the particular element

$${}^{16}_{8}O^{-}_{2}$$

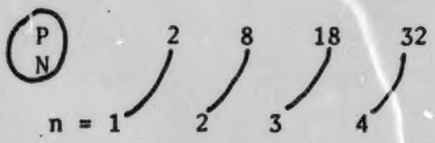
f. Ions

(1) positive Na^+ -loss of e^-

(2) negative O^{-2} = gain of e^-

g. Orbital propellations

$2n^2$ = max possible number e^- in a shell



h. New orbitals - start new orbital each time a new row on the table starts

II Activity

A. Divide the class in groups of four

B. Draw orbital pictures of the first 20 elements

1. Use front cover of text

2. Determine number of neutrons for each of the first 20 elements

3. Page 56, answer items 6, 7, 9

a. For number 9

(1) P = 25
(2) N = 48, e^- = 37
(3) e^- = 78

b. Do 2.26; 2.48; 2.49

APPLICATION Students apply understanding gained here to future work in chemistry

SUMMARY	Summarize important points of the lecture
EVALUATION	Competency with the objectives is evaluated during the lecture exams
HOMEWORK	On course outline

TITLE Structure of the Atom and the Periodic Table

LESSON NUMBER 2

TIME 50 minutes

DATE PREPARED January 26, 1981

BEHAVIORAL OBJECTIVES As a result of attending this lecture students should be able to:

I Describe the structures of atoms in terms of electrons, protons, neutrons

II Draw theoretical pictures of atoms based upon atomic numbers

I CONDITIONS

A. Using class notes and textbooks

B. By the end of class

C. Given the atomic number

II CRITERIA The performance students must demonstrate as an indication they have attained minimum competence with the objectives is:

A. Draw the theoretical structures of the first 18 elements

REFERENCES

I INSTRUCTOR

A. Brown & LeMay sections 2.6; 2.7 and CH 20, pp. 592-3

B. Solutions Guide to Brown & LeMay

II STUDENTS

A. Brown & LeMay

B. Students Guide to Brown & LeMay, Chapter 2

MATERIALS/ EQUIPMENT

I INSTRUCTOR

A. Chalk

B. Chalkboard

II STUDENTS

A. Paper

$1^{o}A = 1 \times 10^{-10}M$
DIA OF ATOM

INTRODUCTION Today we shall begin a discussion of the periodic law and its relationship to atomic theory

INTRODUCTION

This lecture covers many introductory concepts

I I will briefly review the reading material

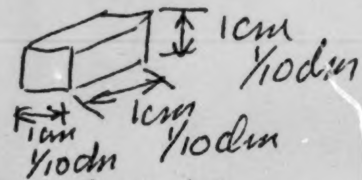
II You will solve problems related to the reading

PROCEDURE

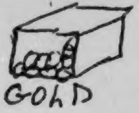
I LECTURE

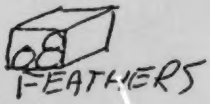
A. Measurements

1. Length - meters
   - a. subdivisions 1/10 dm, cm, mm
   - b. multiples x 10 Km
2. Volume
   - a. length x width x heighth
   - b. $\pi r^2 h$ OR $\pi \frac{d^2}{4}$
   - c. $1mh = 1cm^3$



3. Mass - Kg (1g=1/1000 Kg)- the amount of material
4. Density $D = M/V$





5. Temperature
   - a. Celsius - $C = 9/5(F^{o}-32)$
   - b. Fahrenheit - $F = (5/9\ C) + 32$
   - c. Kelvin - $K = C + 273.15$

     $K = \left[5/9(F-32)\right] + 273.15$

     $F = 32 + \left[9/5(K-273.15)\right]$.
6. Specific Heat
   - a. rate of $\frac{\text{heat capacity of a subst.}}{\text{heat capacity of } H_2O}$
   - b. $J/^{o}C\text{-}g$ $H_2O = 1$ at $15^{o}C$
7. Specific Gravity

B. Dimensional analysis

1. Unit conversion factors to get desired units
   - a. concept of one a/a, 2/2

     $a/b = \frac{1Kg}{2.205\ lbs.}$ or $\frac{2.205\ lbs.}{1Kg}$

2. Examples, P 21 text

C. Classification of Matter

1. Elements, compounds, mixtures

a. Physical changes - ones not involving creation of new substances

b. Chemical changes - conversion of one substance to another

c. Physical properties - do not involve change in the chemical identity of a substance

(1) color
(2) MP
(3) BP
(4) density
(5) solubility
(6) state (solid, liquid, gas)

d. Chemical properties - describe composition and reactivity

(1) use instructor background sheets 1-4

II ACTIVITY IN CLASS

A. Divide class into groups of four students

B. Have students complete sheet #4 using information from sheet #2

C. Do the following problems: 1.34; 1.39; 1.36; 2.4; 2.44 a,b,d

APPLICATION Students apply understanding gained here to their laboratory work and future studies in chemistry

SUMMARY Briefly review the lecture material

EVALUATION Objectives are evaluated in exam #1

HOMEWORK Problems 1.35; 1.38; 1.40; 1.48

LESSON 2

STUDENT ASSIGNMENT SHEET

DIRECTIONS: The reading assignment outlined below should be completed prior to attending Lecture 3. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 2 problem set (listed on separate pages) must be completed and submitted at the beginning of Lecture 4.

READING ASSIGNMENT:

1. Brown and LeMay Chemistry: The Central Science
    - A. Chapter 1, Sections 1.3; 1.5
    - B. Chapter 2, Sections 2.1; 2.2; 2.3; 2.4
    - C. Chapter 4, Section 4.6
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science

PROBLEMS:

1. Chapter 1; 1.35; 1.38; 1.40; 1.48
2. Chapter 2; 2.4; 2.44 a, b, d

TITLE Periodic Law and the Structure of the Atom

LESSON NUMBER 3

TIME 50 Minutes

DATE PREPARED February 2, 1981

BEHAVIORAL OBJECTIVES

As a result of attending this lecture and completing the associated activities, students should be able to:

I Define isotopes
II Write orbital for the first 54 elements
III Draw J and P subshells
IV Relate suborbitals to the structure of the atom

I CONDITIONS

A. Using class notes
B. By the end of the class period

II CRITERIA

The criteria reflecting minimum competence with the objectives are:

A. 90% of all objective related questions must be answered correctly.

REFERENCES

I INSTRUCTOR

A. Brown & LeMay p. 43, 592; Ch 2 section 2.6; Chapter5, p. 136-141; Chapter 6, sections 6.1, 6.2, 6.3, 6.4.

B. Solutions to Exercises in Brown and LeMay.

II STUDENTS

A. "A" above
B. Solutions to Exercises in Brown & LeMay Problems 2.24, 2.28, 6.11, 6.14, 6.40, 6.7.

MATERIALS/ EQUIPMENT

I INSTRUCTOR

A. Chalk
B. Chalkboard
C. One copy of each of the following handouts for each student:

1. Energy Level Diagram for Atomic Orbitals
2. Filling Order of Subshells
3. Ground State Electronic Configuations of Period 2 Atoms Showing Electronic Spins
4. Outer Orbital Formulas
5. Periodic Table

STUDENT

A. Paper
B. One copy each of the handouts listed above
C. Periodic Table

INTRODUCTION

If we are to truly understand bonding relationships we must understand how electrons are placed or situated around the atom. In this lecture, we continue our discussion of the relationship of the atom's structure to the periodic table.

PROCEDURE

I LECTURE

A. Isotopes - atoms of the same element containing different numbers of neutrons:

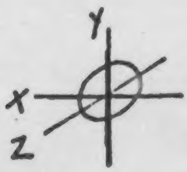
1. ${}^{12}_{6}C$, ${}^{14}_{6}C$
2. ${}^{1}_{1}H$, ${}^{2}_{1}H$, ${}^{3}_{1}H$
3. ${}^{233}_{92}U$, ${}^{236}_{92}U$, ${}^{238}_{92}U$

B. Quantum numbers - orbitals and suborbitals

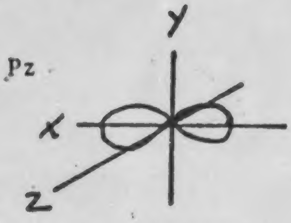
1. n- Principle quantum number
   - a. distance from the nucleus (only)
   - b. related to shells K, L, M, N, O possess several suborbitals.
   - c. EXAMPLE: $\frac{n = 1, 2, 3, 4, 5 \ldots}{\text{Shell}=K,L,M,N,O \ldots} \quad \infty$
2. 1- Azimuthal quantum number
   - a. shape of orbital
   - b. 11 = n-1
   - c. EXAMPLE:

n	1
1	0
2	0, 1
3	0, 1, 2
4	0, 1, 2, 3
	s, p, d, f . . . Shapes

   - d. Suborbital



e. p suborbital Px, Py, Pz



3. me- Magnetic or Orientational quantum number:

   a. describes orbital orientation in space.

   b. me = ±1

   c. EXAMPLE:

1	me	
0	0	(one orbital shape, one orientation)
1	-1, 0, +1	(orbits oriented in 3 directions)
2	-2,-1,0,+1,+2	(orbits oriented in 5 cirections)

4. ms = Spin quantum number

   a. describes the direction of electron spin (clockwise or counterclockwise)

   b. takes on values (-½ & +½)

5. Suborbital populations (s,p,d,f,g,h ...)

SHELL	# OF ELECTRONS ($2n^2$)	ORBITAL LOCATIONS
K(n=1)	2	1s
L(n=2)	8	2s 2p
M(n=3)	8	3s 3p
N(n=4)	32	4s 4p 3d

   a. maximum population of suborbital types

      (1) S = 2
      (2) P = 6
      (3) d = 10
      (4) f = 14

   b. representation of populations in suborbitals

      (1) first period → $1S^2$ ← 2 electrons

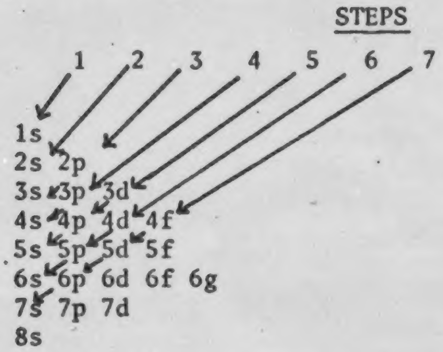
      (2) 2nd period → $2p^6$ ← 6 electrons
      ↖ p shaped orbital

      (3) third period → $3d^{10}$ ← 10 electrons
      ↑ d shaped orbital

6. Shell orbital designation and quantum numbers.

SHELL	QUANTUM NUMBERS				ORBITAL NUMBERS
	n	1	mc	ms	
K	1	0	0	-½,+½	1s
L	2	0	0	-½,+½	2s
	2	1	-1,0,+1	-½,+½	2Py,Py,Pz
M	3	0	0	-½,+½	3s
	3	1	-1,0,+1	-½,+½	3Px,3Py,3Pz
		2	-2,-1,0,1,2	-½,+½	3dz^2, 3dxy,3dxz, 3dyz

7. Orbital filling order



II ACTIVITY - have students begin solving problems listed in in the application section

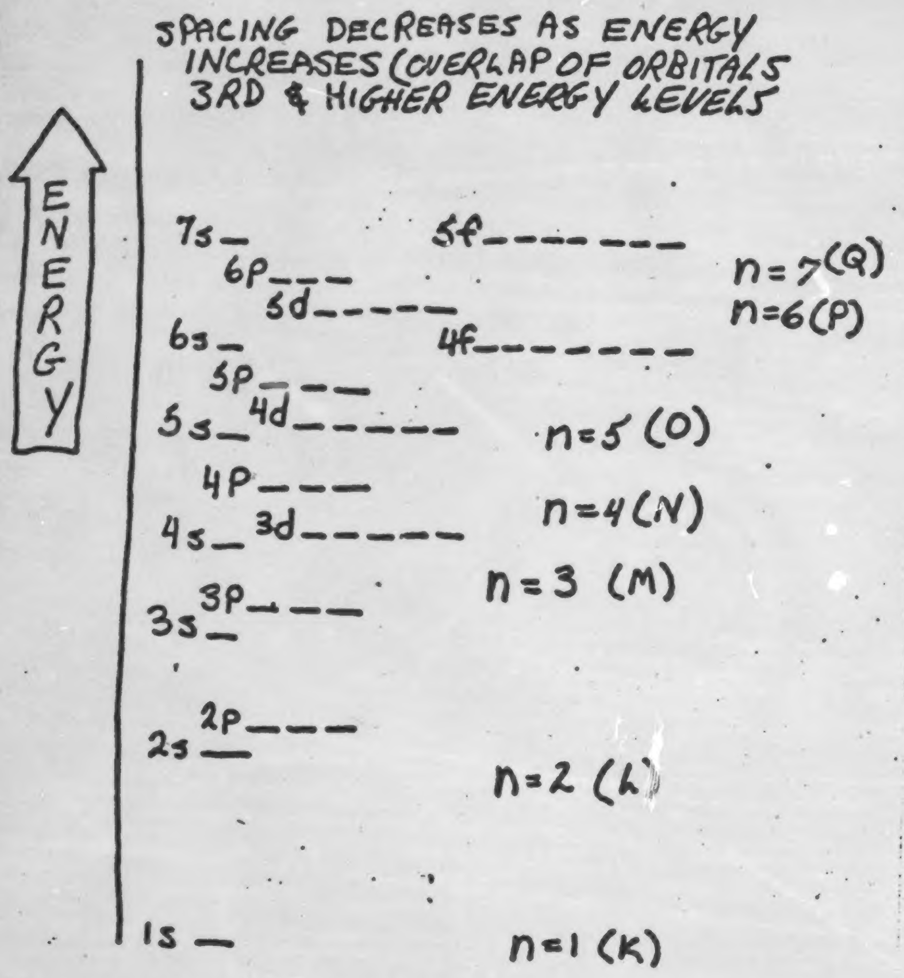
APPLICATION Students apply skills gained here to solving the following problems: 2.24, 2.28, 6.11, 6.14, 6.40, 6.7

SUMMARY Briefly review the mwterial covered in the lecture.

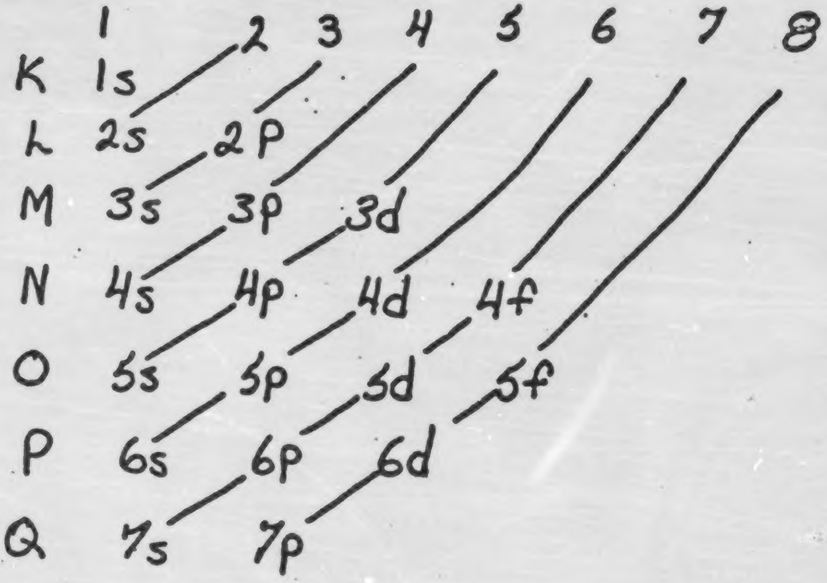
EVALUATION Students are evaluated to determine whether they can demonstrate minimum competency with the objectives in Exam #1.

HOMEWORK Complete work on problems started above.

ENERGY LEVEL DIAGRAM FOR ATOMIC ORBITALS


SPACING DECREASES AS ENERGY INCREASES (OVERLAP OF ORBITALS 3RD & HIGHER ENERGY LEVELS
ENERGY
7s
5f
6p
5d
6s
4f
n=7(Q)
n=6(P)
5p
5s
4d
n=5 (O)
4p
4s
3d
n=4(N)
3p
3s
n=3 (M)
2p
2s
n=2 (L)
1s
n=1 (K)

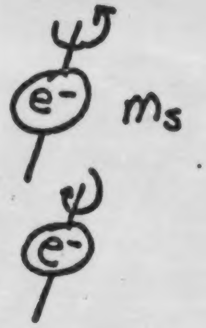
FILLING ORDER OF SUBSHELLS


1
2
3
4
5
6
7
8
K 1s
L 2s 2p
M 3s 3p 3d
N 4s 4p 4d 4f
O 5s 5p 5d 5f
P 6s 6p 6d
Q 7s 7p

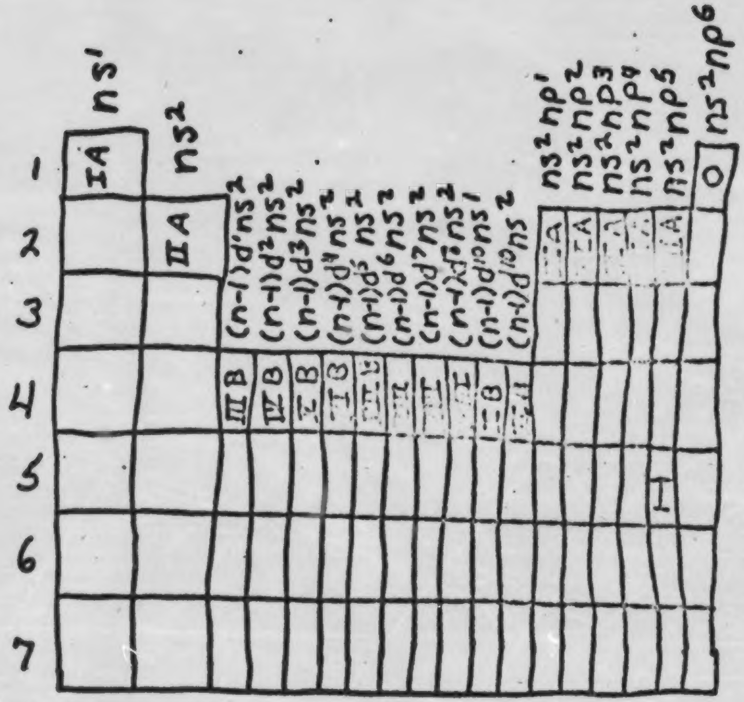
THE NUMERALS AT THE TOP INDICATE THE SEQUENCE THAT IS FOLLOWED IN FILLING SUBSHELLS. SUBSHELLS ON DIAGONAL 2 ARE FILLED BEFORE THOSE ON DIAGONAL 3 WHICH ARE FILLED BEFORE THOSE ON DIAGONAL 4 AND SO ON.

GROUND STATE ELECTRONIC CONFIGURATIONS OF PERIOD 2 ATOMS SHOWING ELECTRONIC SPINS

	1s	2s	$2P_x$	$2P_y$	$2P_z$
Li	↑↓	↑			
Be	↑↓	↑↓			
B	↑↓	↑↓	↑		
C	↑↓	↑↓	↑	↑	
N	↑↓	↑↓	↑	↑	↑
O	↑↓	↑↓	↑↓	↑	↑
F	↑↓	↑↓	↑↓	↑↓	↑
Ne	↑↓	↑↓	↑↓	↑↓	↑↓



OUTER ORBITAL FORMULAS



ns^1
ns^2
IA
IIA
$(n-1)d^1ns^2$
$(n-1)d^2ns^2$
$(n-1)d^3ns^2$
$(n-1)d^4ns^2$
$(n-1)d^5ns^1$
$(n-1)d^6ns^2$
$(n-1)d^7ns^2$
$(n-1)d^8ns^2$
$(n-1)d^{10}ns^1$
$(n-1)d^{10}ns^2$
IIIB
IVB
VB
ns^2np^1
ns^2np^2
ns^2np^3
ns^2np^4
ns^2np^5
ns^2np^6
O
I
1
2
3
4
5
6
7

PERIODS ⟶

FAMOR GROUPS ↓

$I_{\text{OUTER ORBITAL}} = ns^2np^5$

n = PRINCIPLE QUANTUM

$I_{00} = 5s^2 5p^5$

(EXCLUDING LANTHANIDE AND ACTINIDE SERIES)

LESSON 3

SOME FUNDAMENTAL CONCEPTS OF CHEMISTRY

STUDENT ASSIGNMENT SHEET 1

DIRECTIONS Solve the problems listed below using concepts presented in Lecture 3 and the reading assigned for this lesson. Problems should be set up exactly as was done with the model problems in your class text. These problems must be submitted to your instructor at the beginning of Lecture 4.

1. If we assume that density of distilled marine engineering plant water at 4°C to be 1g/ml, at 90°C to be 0.97g/ml and at 4°C 1cm^3 = 1 ml, how much additional volume will be required by 10 liters of water if its temperature increases from 4°C to 90°C ?

2. Feedwater is taken from the sea at 35°F. During the distillation process the H_2O is changed to steam whose temperature rises to at least 105°F. At what centigrade temperature is feedwater taken from the sea and to what Kelvin temperature does the steam temperature rise?

3. One ingredient of corrosion process is dissolved oxygen. Since its solubility in H_2O decreases with increasing temperature steam plant H_2O is often processed through a deceleration heater where the water is heated. If the water temperature is increased here from 280°K to 90°C, what are the starting and ending temperatures in °F ?

4. Hydrocarbon fuels (contain hydrogen and cargon; e.g. gasoline, gunker C, etc.) may be purchased by mass or volume. If bunker C is purchased for shipboard use by volume at 60°F (and its density at this temperature is 0.672g/ml).

   a. At what temperature in °K is the product purchased?

   b. If the temperature at purchase time has increased such that the

the density is 0.596g/ml, how many fewer pounds per gallon are there at the higher temperature?

c. How many fewer Kg will be pumped into your 10,000 gallon starboard fuel tank at the higher than the lower temperature?

d. As chief engineer, you purchase bunker C from Exxon International at $1.00/gallon. How much lower is your bill at the higher temperature?

5. Using dimensional analysis, show how you would convert cubic yards of fuel oil to decimeters.

6. What is the volume of your fuel oil day tank in meters if its length is 20 feet and its diameter is 7 ft?

LESSON 3

STUDENT ASSIGNMENT SHEET

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 4. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 2 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 5.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
   - A. Chapter 2, Sections 2.6; 2.7
   - B. Chapter 20, Introcution, pp. 592-3
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science

PROBLEMS

1. Chapter 2, p. 56, Learning Goals, answer items 6, 7, 9
2. Chapter 2; 2.26; 2.48; 2.49

TITLE	Periodic Law and the Structure of the Atom II
LESSON NUMBER	4
TIME	50 minutes
DATE PREPARED	February 3, 1981
BEHAVIORAL OBJECTIVES	As a result of participation in this lecture and completion of the associated activities, students should be able to: I Predict the suborbital filling order for the majority of elements on the periodic table II Relate suborbital filling order to energy levels III Define Valence electrons IV Define the octet rule and its relationship to molecular bonding theory
I CONDITIONS	The conditions under which students must demonstrate minimum competency with the objectives are: A. Using class notes B. Using class handouts C. By the end of the class period
II CRITERIA	The level reflecting a demonstration of minimum competency with the objectives is: A. Students must answer correctly 90% of all objective related questions
REFERENCES	
I INSTRUCTOR	A. Brown & LeMay - Some readings as lecture 3 B. Brown & LeMay 2.7, 2.8, 7.1, 7.7 C. Solutions to Exercises in Brown and LeMay
II STUDENTS	A. Same as A Above
MATERIALS/ EQUIPMENT	
I INSTRUCTOR	A. Chalk B. Chalkborad C. Student handouts from lecture 3 D. Periodic table
II STUDENTS	A. "C" and "D" above B. Paper

INTRODUCTION

This session is a continuation of the lecture 3. It is important that students understand the relationship between quantum theory and chemistry

PROCEDURES

I LECTURE

A. Review material presented in lecture 3.

1. n,l,me,ms quantum numbers
2. K,L,M,N shells
3. Suborbitals s,p,d,f

B. Discuss the lesson 3 handouts

II ACTIVITY

1. Divide students in groups of four
2. Have each group of students write orbital formulas for the first 54 elements
3. Have students begin work on the following problems 5.50, 5.51, 5.52, 5.30, 5.34

APPLICATION

Students apply background gained here to bonding theory as they continue their study of chemistry

SUMMARY

At the end of the session, review the relationship between the principle quantum number n, the suborbitals s,p,d and f and the periodic table

EVALUATION

Students are evaluated to determine whether they can demonstrate minimum competency with the objectives of this lesson by reviewing the results of exam 1.

HOMEWORK

I. Complete work on problems started above

II. Do the reading outlined on the course syllabus for lecture 5.

LESSON 4

STUDENT ASSIGNMENT SHEET

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 5. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 2 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 6.

READING ASSIGNMENT

1. Brown and LeMay, Chemistry: The Central Science
   - A. p. 43; 592
   - B. Chapter 2, Section 2.6
   - C. Chapter 6, Sections 6.1; 6.2; 6.3; 6.4
   - D. Chapter 5, pp. 136-141 Quantum Numbers and Orbitals
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science
   - A. Chapter 6

PROBLEMS 1. Chapter 2; 2.24; 2.28; 6.11; 6.14; 6.40; 6.7

TITLE

The Relationship of Periodic Classification of
the Elements to Properties of Those Elements

LESSON NUMBER

5

TIME

50 minutes

DATE PREPARED

February 9, 1981

BEHAVIORAL OBJECTIVES

As a result of participation in this lecture and completion of the associated activities students should be able to describe:

I Ways in which elements change as one proceeds from left to right across a period.

II Differences between elements on the left and right sides of the periodic table.

III Ways in which elements change as one proceeds from from top to bottom of a group or family.

IV Uses of the periodic table.

I CONDITIONS

The conditions under which students must demonstrate minimum competency with the objective are:

A. using class notes
B. using class handouts
C. by the end of the class period
D. after having solved the problems listed below for this lesson

II CRITERIA

The level reflecting a demonstration of minimum competency with objectives of this lesson is:

A. Students must answer correctly 90% of all objective related questions when exposed to them in a test situation.

REFERENCES

I INSTRUCTOR

A. Brown & LeMay

1. pp. 45, 133 (ionization energy), 160
2. Chapter 6, section 6.8
3. Chapter 21, section 21.1

B. Solutions to Exercises in Brown & LeMay

II STUDENTS

A. Same as "A" above

B. Hill, chapters 5 and 21

MATERIALS/EQUIPMENT

I INSTRUCTOR

A. Chalk
B. Chalkboard
C. Periodic table
D. Student handouts - one each of the following handouts for each student:

  1. An Introduction to the Periodic Law for General Chemistry Students

II STUDENTS

A. Paper
B. Both C & D above

INTRODUCTION

The periodic table as the name implies contains the elements arranged in a regular fashion. That arrangement allows us to make many predictions about sizes of elements, their ionization potentials, their reactivity, etc. We will consider these as well as uses of the periodic table during this lesson.

PROCEDURE

I. LECTURE

A. Discuss the following from the student handout (also refer to figure 21.3, p. 629, Brown & LeMay)

  1. A breakdown of the periodic table by groups

  2. Valence ring electron populations by period

  3. Electron configurations of the elements

  4. Electron configurations of the elements by groups

  5. Ionization energies by family for the first three periods

  6. Atomic and ionic radii by group

  7. Atomic and ionic radii by period

B. Discuss uses of the periodic table

  1. Bonding relationships

II ACTIVITY

A. Divide class in groups of four students each and have each group:

  1. Identify periodic trends on each of the tables in the student handout.

  2. Have students start to solve problems 5.15, 5.32, 5.35, 5.57, 21.1, 21.14

APPLICATION	Students apply the background gained in this lesson to Lesson 6
SUMMARY	At the conclusion of this session, review the trends in and the uses of the periodic table
EVALUATION	Students are evaluated to determine whether they can demonstrate minimum competency with the objectives of this lesson by reviewing the results of exam 1.
HOMEWORK	I. Complete work on the problems started above II. Do the reading outlined on the course syllabus for Lecture 6

AN INTRODUCTION TO THE PERIODIC LAW FOR GENERAL CHEMISTRY STUDENTS

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INTRODUCTION

When the elements are arranged in the order of increasing atomic number, similarities in physical and chemical properties are repeated at regular intervals. The arrangement implied in this statement is the periodic table (Figure 1).

The table, as most chemistry students soon come to know, is divided horizontally in periods and vertically in groups or families. Accordingly, there are seven periods plus the lanthanide and actinide series, seven groups or families plus the noble gases (sometimes called the inert or nonreactive gases). As the law states, the table is arranged in a predictable manner. It is this manner of arrangement with which we concern ourselves here.

THE TABLE

In the table (Figure 1) elements are arranged horizontally according to their atomic numbers. Each successive element has an atomic number - one greater than its predecessor. Since the atomic numbers indicate the numbers of protons each element has, it can be said that the periodic table is arranged horizontally by increasing atomic mass. Each time the outer most energy level of an element has eight electrons, the maximum extension of the table has been reached in the right-hand direction. Therefore, the right end of each period is ended with one of the nonreactive gases.

A new period always begins at the left-hand edge of the table and the first element in the period always has an atomic number - one greater than the nonreactive gas in the period directly above it. Because atomic numbers of the elements increase from the top to the bottom of the table, as well as from left to right, elements of greatest atomic mass should be found in the lower right-hand corner of the table. Most chemistry texts leave students at this point. From here on students are left to extrapolate the additional reasons for such a periodic arrangement and, subsequently, the types of phenomena this arrangement creates.

Many of the reoccurring relationships which exist as a result of this elemental arrangement must be abstracted from handbooks of chemistry and physics. Others come from relating element electron populations, valence ring electron populations and the like, to the periodic table arrangement.

In an effort to facilitate student understanding of the periodic relationship among the elements, seven tables were constructed in a manner such that periodic trends might be evaluated, both from a periodic and a group or family perspective.

Table 1 provides a breakdown of the periodic table by groups showing the electron populations of each orbital of each element. Included in this table, also, are the ionic or oxidation states each of the elements may assume. Additionally, an asterisk is included after the electron population in each valence ring.

An evaluation of Table 1 shows several interesting reoccurring trends. An element, generally speaking, has the ability to assume at least one ionic state equal to the value of the family number. Again, generally speaking, the electron population of the valence orbital is equal to the family number. The elements in periods 5A, 6A, and 7A assume at least one ionic state equal to the negative value of the difference between 8 and the family number.

8 - FAMILY 6 = -2

Elements in the inert gas family generally do not assume ionic states. These elements all have eight electrons in their outer most or valence shell or energy level. These gases obtain their very stability or nonreactiveness because they have eight electrons in their outermost ring (He is the only exception - it has 2 electrons in its valence ring).

In Table 2 valence ring populations of the various elements have been arranged by period. Each successive element (as one proceeds from left to right across the periodic table) in the A-lettered families has one more valence electron than its predecessor. In each case a period has reached its maximum right hand extension when, as mentioned above, there are eight electrons in the outermost ring.

Table 3 presents electron distribution in the s, p, d, and f suborbitals. From this table it is easy to deduce that the maximum numbers of electrons which may be in s, p, d, and f suborbitals respectively are 2, 6, 10, and 14.

Table 4 presents the electron configurations of each of the elements according to the group in which they exist. From this table it is possible to see the energy levels (the periods) of all electrons belonging to each element. The energy levels are the coefficients of the s, p, d, and f orbitals. The electron population of each element may also be determined from this table. These populations are found by adding the superscripts of each the s, p, d, and f orbitals of each element.

Table 5 provides the ionization energies by family for the first three periods of the periodic table. From the data presented in this table a tremendous increase in ionization energy/mole is noted with the ionization constant following the number of the family in which the element resides. Also, generally speaking, there is a decrease in ionization energy as one proceeds from the top to the bottom of the periodic table.

Tables 6 and 7 consider atomic and ionic radii by group and period respectively. As might be predicted both atomic and ionic radii increase as one proceeds from top to bottom of the periodic table.

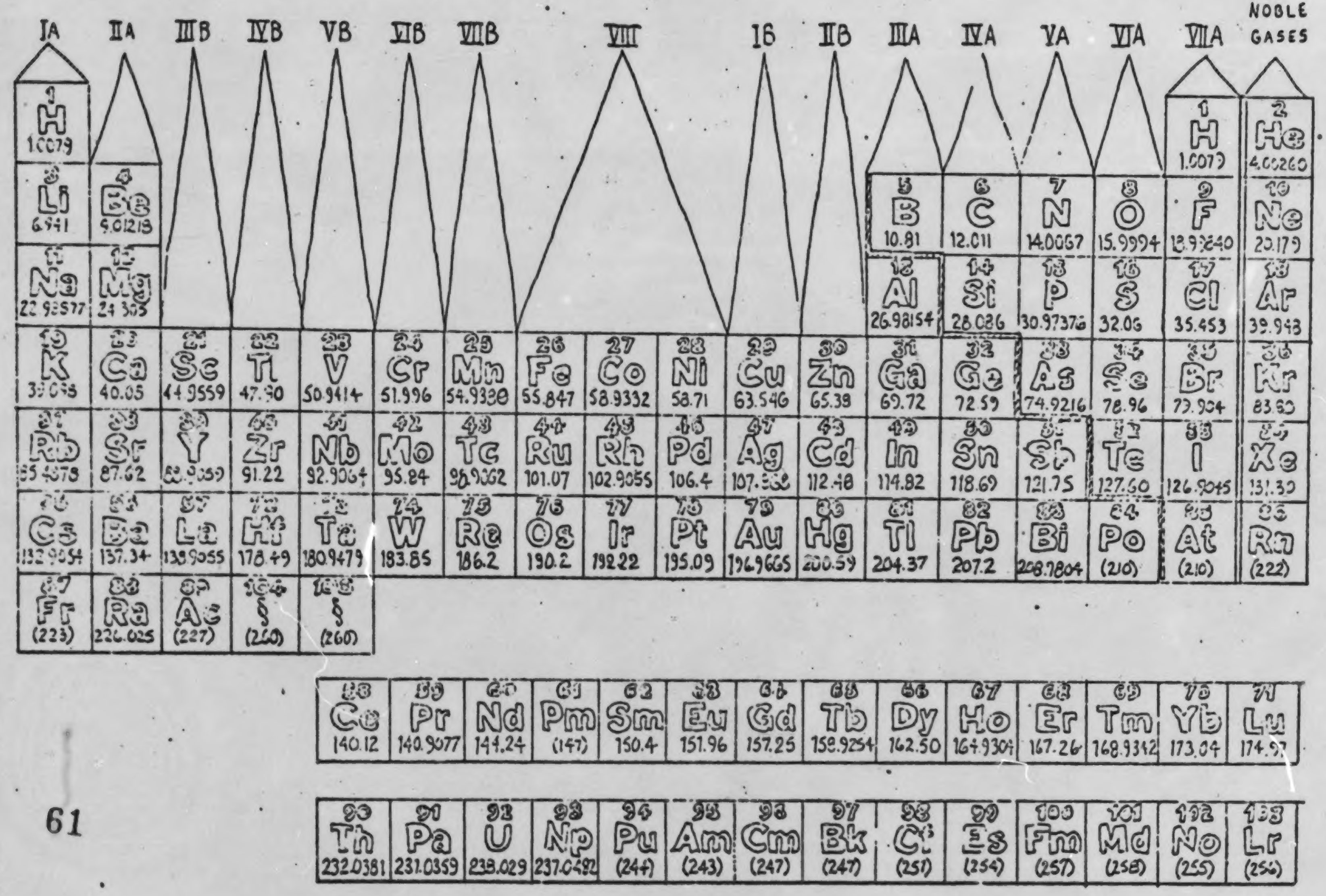
An evaluation of radii trends by period show that as the magnitude of the positive oxidation state increases a decrease in radii occurs. The converse of this trend occurs as the magnitude of the negative oxidation number increases.

CONCLUSIONS

While all identifiable trends have not been discussed, across and down the periodic table, many of the most visible trends have. It is possible to locate and evaluate all trends found throughout the periodic table by dissecting a handbook of chemistry and physics.

- FIGURE #1 -

Periodic Chart of the Elements



IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	NOBLE GASES
1 H 1.0079																1 H 1.0079	2 He 4.00260
3 Li 6.941	4 Be 9.01218											5 B 10.81	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.99840	10 Ne 20.179
11 Na 22.98977	12 Mg 24.305											13 Al 26.98154	14 Si 28.086	15 P 30.97376	16 S 32.06	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.08	21 Sc 44.9559	22 Ti 47.90	23 V 50.9414	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.71	29 Cu 63.546	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.22	41 Nb 92.9064	42 Mo 95.94	43 Tc 98.9062	44 Ru 101.07	45 Rh 102.9055	46 Pd 106.4	47 Ag 107.868	48 Cd 112.40	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.9045	54 Xe 131.30
55 Cs 132.9054	56 Ba 137.34	57 La 138.9055	72 Hf 178.49	73 Ta 180.9479	74 W 183.85	75 Re 186.2	76 Os 190.2	77 Ir 192.22	78 Pt 195.09	79 Au 196.9665	80 Hg 200.59	81 Tl 204.37	82 Pb 207.2	83 Bi 208.9804	84 Po (210)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.025	89 Ac (227)	104 § (260)	105 § (260)													

58 Ce 140.12	59 Pr 140.9077	60 Nd 144.24	61 Pm (147)	62 Sm 150.4	63 Eu 151.96	64 Gd 157.25	65 Tb 158.9254	66 Dy 162.50	67 Ho 164.9304	68 Er 167.26	69 Tm 168.9342	70 Yb 173.04	71 Lu 174.97
90 Th 232.0381	91 Pa 231.0359	92 U 238.029	93 Np 237.0482	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (254)	100 Fm (257)	101 Md (258)	102 No (255)	103 Lr (256)

TABLE 1

A BREAKDOWN OF THE PERIODIC TABLE BY GROUPS OR FAMILIES

FAMILY	PERIOD	ELEMENT SYMBOL	ELEMENT NUMBER	COMMON IONIC STATE/STATES	ELECTRON POPULATIONS BY ORBITS K 1	L 2	M 3	N 4	O 5	P 6	Q 7
1A	1	H	1	1^+	1*						
1A	2	Li	3	1^+	2	1*					
1A	3	Na	11	1^+	2	8	1*				
1A	4	K	19	1^+	2	8	8	1*			
1A	5	Rb	37	1^+	2	8	18	8	1*		
1A	6	Cs	55	1^+	2	8	18	18	8	1*	
1A	7	Fr	87	1^+	2	8	18	32	18	8	1*
2A	2	Be	4	2^+	2	2*					
2A	3	Mg	12	2^+	2	8	2*				
2A	4	Ca	20	2^+	2	8	8	2*			
2A	5	Sr	38	2^+	2	8	18	8	2*		
2A	6	Ba	56	2^+	2	8	18	18	8	2*	
[illegible]	7	Ra	88	2^+	2	8	18	32	18	8	2*
[illegible]	2	B	5	3^+	2	3*					
3A	3	Al	13	3^+	2	8	3*				
3A	4	Ga	31	3^+	2	8	10	8	3*		
3A	5	In	49	3^+	2	8	18	10	8	3*	
3A	6	Tl	81	1^+, 3^+	2	8	18	32	10	8	3*
4A	2	C	6	2^+, 4^+, 4^-	2	4*					
4A	3	Si	14	2^+, 4^+, 4^-	2	8	4*				
4A	4	Ge	32	2^+, 4^+	2	8	10	8	4*		
4A	5	Sn	50	2^+, 4^+	2	8	18	10	8	4*	
4A	6	Pb	82	2^+, 4^+	2	8	18	32	10	8	4*
5A	2	N	7	1^+, 2^+, 3^+, 4^+, 5^+, 1^-, 2^-, 3^-	2	5*					
5A	3	P	15	2^+, 4^+, 4^-	2	8	5*				
5A	4	As	33	3^+, 5^+, 3^-	2	8	10	8	5*		
5A	5	Sb	51	3^+, 5^+, 3^-	2	8	18	10	8	5*	
5A	6	Bi	83	3^+, 5^+	2	8	18	32	10	8	5*
6A	2	O	8	2^-	2	6*					
6A	3	S	16	4^+, 6^+, 2^-	2	8	6*				
6A	4	Se	34	4^+, 6^+, 2^-	2	8	10	8	6*		
6A	5	Te	52	4^+, 6^+, 2^-	2	8	18	10	8	6*	
6A	6	Po	84	2^+, 4^+	2	8	18	32	10	8	6*
7A	1	F	9	1^-	2	7*					
7A	3	Cl	17	1^+, 5^+, 7^+, 1^-	2	8	7*				
7A	4	Br	35	1^+, 5^+, 1^-	2	8	10	8	7*		
7A	5	I	53	1^+, 5^+, 7^+, 1^-	2	8	18	10	8	7*	
7A	6	At	85	–	2	8	18	32	10	8	7*
0	1	He	2	0	2*						
0	2	Ne	10	0	2	8*					
0	3	Ar	18	0	2	8	8*				
0	4	Kr	36	0	2	8	18	8*			
0	5	Xe	54	0	2	8	18	18	8*		
0	6	Rn	86	0	2	8	18	32	10	8	8*
3B	4	Sc	21	3^+	2	8	9	2*			
3B	5	Y	39	3^+	2	8	18	9	2*		
3B	6	La	57	3^+	2	8	18	18	9	2*	
3B	7	Ac	89	3^+	2	8	18	32	18	9	2*
4B	4	Ti	22	2^+, 3^+, 4^+	2	8	10	2*			
4B	5	Zr	40	4^+	2	8	18	10	2*		
4B	6	Hf	72	4^+	2	8	18	32	10	2*	
4B	7	?	104	4^+	2	8	18	32	32	10	2*
5B	4	V	23	2^+, 3^+, 4^+, 5^+	2	8	11	2*			
5B	5	Nb	41	3^+, 5^+	2	8	18	12	1*		
5B	6	Ta	73	5^+	2	8	18	32	11	2*	
5B	7	?	105	–	2	8	18	32	32	11	2*
6B	4	Cr	24	2^+, 3^+, 6^+	2	8	13	1*			
6B	5	Mo	42	6^+	2	8	18	13	1*		
6B	6	W	74	6^+	2	8	18	32	13	1*	
6B	7	?	106	–	–						
7B	4	Mn	25	2^+, 3^+, 4^+, 7^+	2	8	13	2*			
7B	5	Tc	43	4^+, 6^+, 7^+	2	8	18	13	2*		
7B	6	Re	75	4^+, 6^+, 7^+	2	8	18	32	13	2*	

TABLE I (CONTINUED)

A BREAKDOWN OF THE PERIODIC TABLE BY GROUPS OR FAMILIES

FAMILY	PERIOD	ELEMENT SYMBOL	ELEMENT NUMBER	COMMON IONIC STATE/STATES	ELECTRON POPULATIONS BY ORBITS						
					K 1	L 2	M 3	N 4	O 5	P 6	Q 7
8	8	Fe	26	2^+, 3^+	2	8	14	2*			
8	5	Ru	44	3^+	2	8	18	14	2*		
8	6	Os	76	3^+, 4^+	2	8	18	32	14	2*	
8	4	Co	27	2^+, 3^+	2	8	15	2*			
8	5	Rh	45	3^+	2	8	18	15	2*		
8	6	Ir	77	3^+, 4^+	2	8	18	32	15	2*	
8	4	Ni	28	2^+, 3^+	2	8	16	2*			
8	5	Pd	46	2^+, 4^+	2	8	18	18*			
8	6	Pt	78	2^+, 4^+	2	8	18	32	16	2*	
1B	4	Cu	29	1^+, 2^+	2	8	18	1*			
1B	5	Ag	47	1^+	2	8	18	18	1*		
1B	6	Au	79	1^+, 3^+	2	8	18	32	18	1*	
2B	4	Zr	30	2^+	2	8	18	2*			
2B	5	Cd	48	2^+	2	8	18	18	2*		
2B	6	Hg	80	1^+, 2^+	2	8	18	32	18	2*	

TABLE 2

VALENCE RING ELECTRON POPULATIONS BY PERIOD

PERIOD	GROUP OR FAMILY	ELEMENT #	ELEMENT	ELECTRON POPULATIONS IN VALENCE RINGS						
				K 1	L 2	M 3	N 4	O 5	P 6	Q 7
1	1A	1	H	1						
1	0	2	He	2						
2	1A	3	Li		1					
2	2A	4	Be		2					
2	3A	5	B		3					
2	4A	6	C		4					
2	5A	7	N		5					
2	6A	8	O		6					
2	7A	9	F		7					
2	0	10	Ne		8					
3	1A	11	Na			1				
3	2A	12	Ma			2				
3	3A	13	Mg			3				
3	4A	14	Si			4				
3	5A	15	P			5				
3	6A	16	S			6				
3	7A	17	Cl			7				
3	0	18	Ar			8				
4	1A	19	K				1			
4	2A	20	Ca				2			
4	3B	21	Sc				2			
4	4B	22	Ti				2			
4	5B	23	V				2			
4	6B	24	Cr				1			
4	7B	25	Mn				2			
4	8	26	Fe				2			
4	8	27	Co				2			
4	8	28	Ni				2			
4	1B	29	Cu				1			
4	2B	30	Zn				2			
4	3A	31	Ga				3			
4	4A	32	Ge				4			
4	5A	33	As				5			
4	6A	34	Se				6			
4	7A	35	Br				7			
4	0	36	Kr				8			
5	1A	37	Rb					1		
5	2A	38	Sr					2		
5	3B	39	Y					2		
5	4B	40	Zr					2		
5	5B	41	Nb					1		
5	6B	42	Mo					1		
5	7B	43	Tc					2		
5	8	44	Ru					1		
5	8	45	Rh					1		
5	8	46	Pd					18		
5	1B	47	Ag					1		
5	2B	48	Cd					2		
5	3A	49	In					3		
5	4A	50	Sn					4		
5	5A	51	Sb					5		
5	6A	52	Te					6		
5	7A	53	I					7		
5	0	54	Xe					8		
6	1A	55	Cs						1	
6	2A	56	Ba						2	
6	3B	57	La						2	
6	4B	72	Hf						2	
6	5B	73	Ta						2	
6	6B	74	W						1	
6	7B	75	Re						2	
6	8	76	Os						2	

TABLE 2 (CONTINUED)

VALENCE RING ELECTRON POLULATIONS BY PERIOD

PERIOD	GROUP OR FAMILY	ELEMENT #	ELEMENT	ELECTRON POPULATIONS IN VALENCE RINGS						
				K 1	L 2	M 3	N 4	O 5	P 6	Q 7
6	8	77	Ir						2	
6	8	78	Pt						2	
6	1B	79	Au						1	
6	2B	80	Hg						2	
6	3A	81	Tl							3
6	4A	82	Pb							4
6	5A	83	Bi							5
6	6A	84	Po							6
6	7A	85	At							7
6	0	86	Rn							8
7	1A	67	Fr							1
7	2A	88	Ra							2
7	3B	89	Ac							3

TABLE 3

ELECTRONIC CONFIGURATIONS OF THE ELEMENTS

ATOMIC NUMBER	ELEMENT	ℓ orbitals																		
		s	s	p	s	p	d	s	p	d	f	s	p	d	f	s	p	d	f	s
1	H	1																		
2	He	2																		
3	Li	2	1																	
4	Be	2	2																	
5	B	2	2	1																
6	C	2	2	2																
7	N	2	2	3																
8	O	2	2	4																
9	F	2	2	5																
10	Ne	2	2	6																
11	Na	2	2	6	1															
12	Mg	2	2	6	2															
13	Al	2	2	6	2	1														
14	Si	2	2	6	2	2														
15	P	2	2	6	2	3														
16	S	2	2	6	2	4														
17	Cl	2	2	6	2	5														
18	Ar	2	2	6	2	6														
19	K	2	2	6	2	6		1												
20	Ca	2	2	6	2	6		2												
21	Sc	2	2	6	2	6	1	2												
22	Ti	2	2	6	2	6	2	2												
23	V	2	2	6	2	6	3	2												
24	Cr	2	2	6	2	6	5^x	1												
25	Mn	2	2	6	2	6	5	2												
26	Fe	2	2	6	2	6	6	2												
27	Co	2	2	6	2	6	7	2												
28	Ni	2	2	6	2	6	8	2												
29	Cu	2	2	6	2	6	10^x	1												
30	Zn	2	2	6	2	6	10	2												
31	Ga	2	2	6	2	6	10	2	1											
32	Ge	2	2	6	2	6	10	2	2											
33	As	2	2	6	2	6	10	2	3											
34	Se	2	2	6	2	6	10	2	4											
35	Br	2	2	6	2	6	10	2	5											
36	Kr	2	2	6	2	6	10	2	6											
37	Rb	2	2	6	2	6	10	2	6			1								
38	Sr	2	2	6	2	6	10	2	6			2								
39	Y	2	2	6	2	6	10	2	6	1		2								
40	Zr	2	2	6	2	6	10	2	6	2		2								
41	Nb	2	2	6	2	6	10	2	6	4^x		1								
42	Mo	2	2	6	2	6	10	2	6	5		1								
43	Tc	2	2	6	2	6	10	2	6	6		1								
44	Ru	2	2	6	2	6	10	2	6	7		1								
45	Rh	2	2	6	2	6	10	2	6	8		1								
46	Pd	2	2	6	2	6	10	2	6	10^x										
47	Ag	2	2	6	2	6	10	2	6	10		1								
48	Cd	2	2	6	2	6	10	2	6	10		2								
49	In	2	2	6	2	6	10	2	6	10		2	1							
50	Sn	2	2	6	2	6	10	2	6	10		2	2							
51	Sb	2	2	6	2	6	10	2	6	10		2	3							
52	Te	2	2	6	2	6	10	2	6	10		2	4							
53	I	2	2	6	2	6	10	2	6	10		2	5							
54	Xe	2	2	6	2	6	10	2	6	10		2	6							
55	Cs	2	2	6	2	6	10	2	6	10		2	6			1				
56	Ba	2	2	6	2	6	10	2	6	10		2	6			2				
57	ha	2	2	6	2	6	10	2	6	10		2	6	1		2				
58	Ce	2	2	6	2	6	10	2	6	10	2^x	2	6			2				
59	Pr	2	2	6	2	6	10	2	6	10	3	2	6			2				
60	Nd	2	2	6	2	6	10	2	6	10	4	2	6			2				
61	Pm	2	2	6	2	6	10	2	6	10	5	2	6			2				
62	Sm	2	2	6	2	6	10	2	6	10	6	2	6			2				
63	Eu	2	2	6	2	6	10	2	6	10	7	2	6			2				
64	Gd	2	2	6	2	6	10	2	6	10	7^x	2	6	1		2				

TABLE 3 (CONTINUED)

ELECTRONIC CONFIGURATIONS OF THE ELEMENTS

ATOMIC NUMBER	ELEMENT	ℓ orbitals																		
		s	s	p	s	p	d	s	p	d	f	s	p	d	f	s	p	d	f	s
65	Tb	2	2	6	2	6	10	2	6	10	9	2	6			2				
66	Dy	2	2	6	2	6	10	2	6	10	10	2	6			2				
67	Ho	2	2	6	2	6	10	2	6	10	11	2	6			2				
68	Er	2	2	6	2	6	10	2	6	10	12	2	6			2				
69	Tm	2	2	6	2	6	10	2	6	10	13	2	6			2				
70	Yb	2	2	6	2	6	10	2	6	10	14	2	6			2				
71	Lu	2	2	6	2	6	10	2	6	10	14	2	6	1		2				
72	Hf	2	2	6	2	6	10	2	6	10	14	2	6	2		2				
73	Ta	2	2	6	2	6	10	2	6	10	14	2	6	3		2				
74	W	2	2	6	2	6	10	2	6	10	14	2	6	4		2				
75	Re	2	2	6	2	6	10	2	6	10	14	2	6	5		2				
76	Os	2	2	6	2	6	10	2	6	10	14	2	6	6		2				
77	Ir	2	2	6	2	6	10	2	6	10	14	2	6	7		2				
78	Pt	2	2	6	2	6	10	2	6	10	14	2	6	9		1				
79	Au	2	2	6	2	6	10	2	6	10	14	2	6	10		1				
80	Hg	2	2	6	2	6	10	2	6	10	14	2	6	10		2				
81	Tl	2	2	6	2	6	10	2	6	10	14	2	6	10		2	1			
82	Pb	2	2	6	2	6	10	2	6	10	14	2	6	10		2	2			
83	Bi	2	2	6	2	6	10	2	6	10	14	2	6	10		2	3			
84	Po	2	2	6	2	6	10	2	6	10	14	2	6	10		2	4			
85	At	2	2	6	2	6	10	2	6	10	14	2	6	10		2	5			
86	Rn	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6			
87	Fr	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6			1
88	Ra	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6			2
89	Ac	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6	1		2
90	Th	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6	2		2
91	Pa	2	2	6	2	6	10	2	6	10	14	2	6	10	2^{x}	2	6	1		2
92	U	2	2	6	2	6	10	2	6	10	14	2	6	10	3	2	6	1		2
93	Np	2	2	6	2	6	10	2	6	10	14	2	6	10	4	2	6	1		2
94	Pu	2	2	6	2	6	10	2	6	10	14	2	6	10	6^{x}	2	6			2
95	Am	2	2	6	2	6	10	2	6	10	14	2	6	10	7	2	6			2
96	Cm	2	2	6	2	6	10	2	6	10	14	2	6	10	7	2	6	1		2
97	Bk	2	2	6	2	6	10	2	6	10	14	2	6	10	9^{x}	2	6			2
98	Cf	2	2	6	2	6	10	2	6	10	14	2	6	10	10	2	6			2
99	Es	2	2	6	2	6	10	2	6	10	14	2	6	10	11	2	6			2
100	Fm	2	2	6	2	6	10	2	6	10	14	2	6	10	12	2	6			2
101	Md	2	2	6	2	6	10	2	6	10	14	2	6	10	13	2	6			2
102	No	2	2	6	2	6	10	2	6	10	14	2	6	10	14	2	6			2
103	Lr	2	2	6	2	6	10	2	6	10	14	2	6	10	14	2	6	1		2
104	Rf	2	2	6	2	6	10	2	6	10	14	2	6	10	14	2	6	2		2

^{x} Does not follow the predicted order

TABLE 4

ELECTRON CONFIGURATIONS OF ELEMENTS BY GROUP

GROUP	ATOMIC NUMBER	ELEMENT	ELECTRON CONFIGURATION
I A	1	H	$1S^1$
I A	3	Li	He + $2s^1$ (He = $1s^2$)
I A	11	Na	Ne + $3s^1$ (Ne = $2s^2 2p^6$)
I A	19	K	Ar + $4s^1$ (Ar = Ne + $3s^2 3p^6$)
I A	37	Rb	Kr + $5s^1$ (Kr = Ar + $4s^2 3d^{10} 4p^6$)
I A	55	Cs	Xe + $6s^1$ (Xe = Kr + $5s^2 4d^{10} 5p^6$)
I A	87	Fr	Rn + $7s^1$ (Rn = Xe + $6s^2 4f^{14} 5d^{10} 6p^6$)
II A	4	Be	He + $2s^2$
II A	12	Mg	Ne + $3s^2$
II A	20	Ca	Ar + $4s^2$
II A	38	Sr	Kr + $5s^2$
II A	56	Ba	Xe + $6s^2$
II A	88	Ra	Rn + $7s^2$
III B	21	Sc	Ar + $4s^2 3d^1$
III B	39	Y	Kr + $5s^2 4d^1$
III B	57	La	Xe + $6s^2 5d^1$
III B	89	Ac	Rn + $7s^2 6d^1$
IV B	22	Ti	Ar + $3s^2 3d^2$
IV B	40	Zr	Kr + $5s^2 4d^2$
IV B	72	Hf	Xe + $6s^2 4f^{14} 5d^2$
V B	23	V	Ar + $4s^2 3d^3$
V B	41	Nb	Kr + $5s^1 4d^4$
V B	73	Ta	Xe + $6s^2 4f^{14} 5d^3$
VI B	24	Cr	Ar + $4s^1 3d^5$
VI B	42	Mo	Kr + $5s^1 4d^5$
VI B	74	W	Xe + $6s^2 4f^{14} 5d^4$
VII B	25	Mn	Ar + $4s^2 3d^5$
VII B	43	Tc	Kr + $5s^2 4d^5$
VII B	75	Re	Xe + $6s^2 4f^{14} 5d^5$
VIII	26	Fe	Ar + $4s^2 3d^6$
VIII	44	Ru	Kr + $5s^1 4d^7$
VIII	76	Os	Xe + $6s^2 4f^{14} 5d^6$
VIII	27	Co	Ar + $4s^2 3d^7$
VIII	45	Rh	Kr + $5s^1 4d^8$
VIII	77	Ir	Xe + $6s^2 4f^{14} 5d^7$
VIII	28	Ni	Ar + $4s^2 3d^8$
VIII	46	Pd	Kr + $4d^{10}$
VIII	78	Pt	Xe + $6s^1 4f^{14} 5d^9$
I B	29	Cu	Ar + $4s^1 3d^{10}$
I B	47	Ag	Kr + $5s^1 4d^{10}$
I B	79	Au	Xe + $6s^1 4f^{14} 5d^{10}$
II B	30	Zn	Ar + $4s^2 3d^{10}$
II B	48	Cd	Kr + $5s^2 4d^{10}$
II B	80	Hg	Xe + $6s^2 4f^{14} 5d^{10}$
III A	5	B	He + $2s^2 2p^1$
III A	13	Al	Ne + $3s^2 3p^1$
III A	31	Ga	Ar + $4s^2 3d^{10} 4p^1$
III A	49	In	Kr + $5s^2 4d^{10} 5p^1$
III A	81	Tl	Xe + $6s^2 4f^{14} 5d^{10} 6p^1$
IV A	6	C	He + $2s^2 2p^2$
IV A	14	Si	Ne + $3s^2 3p^2$
IV A	32	Ge	Ar + $4s^2 3d^{10} 4p^2$
IV A	50	Sn	Kr + $5s^2 4d^{10} 5p^2$
IV A	82	Pb	Xe + $6s^2 4f^{14} 5d^{10} 6p^2$
V A	7	N	He + $2s^2 2p^3$
V A	15	P	Ne + $3s^2 3p^3$
V A	33	As	Ar + $4s^2 3d^{10} 4p^3$
V A	51	Sb	Kr + $5s^2 4d^{10} 5p^3$
V A	83	Bi	Xe + $6s^2 4f^{14} 5d^{10} 6p^3$
VI A	8	O	He + $2s^2 2p^4$
VI A	16	S	Ne + $3s^2 3p^4$
VI A	34	Se	Ar + $4s^2 3d^{10} 4p^4$
VI A	52	Te	Kr + $5s^2 4d^{10} 5p^4$
VI A	84	Po	Xe + $6s^2 4f^{14} 5d^{10} 6p^4$
VII A	9	F	He + $2s^2 2p^5$

TABLE 4 (CONTINUED)

ELECTRON CONFIGURATIONS OF ELEMENTS BY GROUP

GROUP	ATOMIC NUMBER	ELEMENT	ELECTRON CONFIGURATION
VII A	17	Cl	Ne + $3s^2 3p^5$
VII A	35	Br	Ar + $4s^2 3d^{10} 4p^5$
VII A	53	I	Kr + $5s^2 4d^{10} 5p^5$
VII A	85	At	Xe + $6s^2 4f^{14} 5d^{10} 6p^5$
0	2	He	$1S^2$
0	10	Ne	He + $2s^2 2p^6$
0	18	Ar	Ne + $3s^2 3p^6$
0	36	Kr	Ar + $4s^2 3d^{10} 4p^6$
0	54	Xe	Kr + $5s^2 4d^{10} 5p^6$
0	86	Rn	Xe + $6s^2 4f^{14} 5d^{10} 6p^6$

TABLE 5

IONIZATION ENERGIES IN K JOULES/MOLE BY FAMILY FOR THE FIRST THREE PERIODS

FAMILY	ELEMENT	FIRST	SECOND	THIRD	FOURTH	FIFTH	SIXTH	SEVENTH	EIGHTH
I A	H	1312							
I A	Li	520	7297	11810					
I A	Na	495.8	4562	6912	9540	13360	16610	20110	25490
I A	K	418.8	3069	4600	5879	7971	9619	11380	14950
II A	Be	900	1757	14840	21000				
II A	Mg	737.6	1450	7732	10550	13620	18000	21700	25660
II A	Ca	598.5	1146	4941	6485	8142	10520	12350	13830
III A	B	800	2430	3659	25020	32810			
III A	Al	577.4	1816	2744	11580	15030	18370	23290	27460
IV A	C	1086	2352	4619	6221	37800	47300		
IV A	Si	786.2	1577	3229	4356	16080	19790	23780	29250
V A	N	1402	2857	4577	7473	9443	53250	64340	
V A	P	1012	1896	2910	4954	6272	21270	25410	29840
VI A	O	1314	3391	5301	7468	10980	13320	71300	84050
VI A	S	999.6	2260	3380	4565	6996	8490	28080	13720
VII A	F	1681	3375	6045	8418	11020	15160	17860	92000
VII A	Cl	1255	2297	3850	5146	6544	9330	11020	33600
O	He	2371	5247						
O	Ne	2080	3963	6276	9376	12190	15230		
O	Ar	1520	2665	3947	5770	7240	8810	11970	13840

TABLE 6

ATOMIC AND IONIC RADII (IN ANGSTROMS) BY GROUP

GROUP	ELEMENT	PERIOD	ATOMIC RADIUS	IONIC RADIUS	CHARGE	VALENCE E^-
I A	Li	2	1.35	0.60	+ 1	$2s^1$
I A	Na	3	1.54	0.95	+ 1	$3s^1$
I A	K	4	1.96	1.33	+ 1	$4s^1$
I A	Rb	5	2.11	1.48	+ 1	$5s^1$
I A	Cs	6	2.25	1.69	+ 1	$6s^1$
II A	Be	2	0.90	0.31	+ 2	$2s^2$
II A	Mg	3	1.35	0.65	+ 2	$3s^2$
II A	Ca	4	1.74	0.99	+ 2	$4s^2$
II A	Sr	5	1.92	1.13	+ 2	$5s^2$
II A	Ba	6	1.98	1.35	+ 2	$6s^2$
III A	Al	3	1.43	0.50	+ 3	$3s^2 3p^1$
III A	Ga	4	1.22	0.62	+ 3	$4s^2 4p^1$
III A	In	5	1.62	0.81	+ 3	$5s^2 5p^1$
V A	N	2	0.70	1.71	- 3	$2s^2 2p^3$
V A	P	3	1.10	2.12	- 3	$3s^2 3p^3$
VI A	O	2	0.66	1.40	- 2	$2s^2 2p^4$
VI A	S	3	1.04	1.84	- 2	$3s^2 3p^4$
VI A	Se	4	1.17	1.98	- 2	$4s^2 4p^4$
VI A	Te	5	1.37	2.21	- 2	$5s^2 5p^4$
VII A	F	2	0.64	1.36	- 1	$3s^2 3p^5$
VII A	Cl	3	0.99	1.81	- 1	$3s^2 3p^5$
VII A	Br	4	1.14	1.95	- 1	$4s^2 4p^5$
VII A	I	5	1.33	2.16	- 1	$5s^2 5p^5$

TABLE 7

ATOMIC AND IONIC RADII (IN ANGSTROMS) BY PERIOD

PERIOD	ELEMENT	GROUP	ATOMIC NUMBER	ATOMIC RADIUS	IONIC RADIUS	CHARGE
2	Li	I A	3	1.35	0.60	+ 1
2	Be	II A	4	0.90	0.31	+ 2
2	N	V A	5	0.70	1.71	- 3
2	O	VI A	6	0.66	1.40	- 2
2	F	VII A	7	0.64	1.36	- 1
3	Na	I A	11	1.54	0.95	+ 1
3	Mg	II A	12	1.35	0.65	+ 3
3	Al	III A	13	1.43	0.50	+ 3
3	P	V A	15	1.10	2.12	- 3
3	S	VI A	16	1.04	1.84	- 2
3	Cl	VII A	17	0.99	1.81	- 1
4	K	I A	19	1.96	1.33	+ 1
4	Ca	II A	20	1.74	0.99	+ 2
4	Ga	III A	31	1.22	0.62	+ 3
4	Se	VI A	34	1.17	1.98	- 2
4	Br	VII A	35	1.14	1.95	- 1
5	Rb	I A	37	2.11	1.48	+ 1
5	Sr	II A	38	1.92	1.13	+ 2
5	In	III A	49	1.62	0.81	+ 3
5	Te	VI A	52	1.37	2.21	- 2
5	I	VII A	53	1.33	2.16	- 1
6	Cs	I A	55	2.25	1.69	+ 1
6	Ba	II A	56	1.98	1.35	+ 2

LESSON 5

STUDENT ASSIGNMENT SHEET

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 6. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 2 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 7.

READING ASSIGNMENT

1. Brown and LeMay Chemistry; The Central Science
   - A. Chapter 2, Sections 2.7; 2.8
   - B. Chapter 7, Sections 7.1; 7.7

PROBLEMS 1. Chapter 5; 5.30; 5.34; 5.50; 5.51; 5.52

TITLE General Concepts of Chemical Bonding

LESSON NUMBER 6

TIME 50 minutes

DATE PREPARED February 10, 1981

BEHAVIORAL OBJECTIVES

As a result of participation in this lecture and completion of the associated activities, students should be able to:

I Predict molecular formulae using common oxidation numbers

II Describe the difference between ionic and covalent character of bonds

III Determine the degree of ionic character based upon electronegativities

I CONDITIONS

The conditions under which students must demonstrate minimum competency with the objectives are:

A. Using class notes
B. Using class handouts
C. Using class textual materials
D. By the end of the class period
E. After having solved the problems listed below for this lesson

II CRITERIA

The level reflecting a demonstration of minimum competence with the objectives of this lesson is:

A. Students must answer correctly 90% of all objective related questions when exposed to them in a text situation

REFERENCE

I INSTRUCTOR

A. Brown & LeMay
  1. Chapter 2, section 2.8
  2. Chapter 7, sections 7.1, 7.2, 7.4, 7.5, 7.9, 7.10

II STUDENTS

A. Same as "A" above
B. Hill, Chapter 7

MATERIALS/EQUIPMENT

I INSTRUCTOR

A. Chalk
B. Chalkboard
C. Periodic table
D. One copy of student handout 6-1 "Oxidation Numbers and Bond Polarities"

II STUDENTS

A. Paper
B. Periodic table

INTRODUCTION

Today we begin a discussion of bonding theory. The use of oxidation numbers and electronegativities are an integrated part of that discussion.

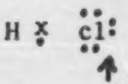
PROCEDURE

I LECTURE

A. Period table
   1. Meaning of oxidation numbers - their relationship to the octet rule.

      $H(1^+)$, $O(2^-)$ $H_2+O \rightarrow H_2O$

   2. Using electron dot (Lewis) structures to represent valence electrons:



valence electrons

   3. Single, double and triple bonds.

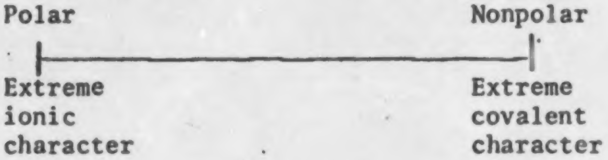
B. Bonds

   1. Ionic bonds - the transfer of one or more electrons from one atom to form a molecule

      a. Molecult-atoms, tightly bound together

   2. Covalent bonds - electrons are shared

   3. Bond continuum



      a. Electronegativity - the tendency of atoms to attract electrons

         (1) The difference between the electronegativities of two bonding atoms determines the degree of bond polarity

$$\underset{2.2}{H} - \underset{3.2}{Cl}$$

$$\underset{2.2}{H} - \underset{4.0}{F}$$

II ACTIVITY

A. Divide class into groups of 4 students each

B. Have students conduct the activity outlined on student handout 6-1, Oxidation Numbers and Bond Polarities.

C. Have students begin work on the following textbook problems: 2.28, 2.32, 2.52a, 2.57, 7.1, 7.2, 7.4, 7.7, 7.10.

APPLICATION

I. Students apply skills and understanding gained here to:

A. Work in lesson 7
B. Future work in chemistry

SUMMARY

At the end of the class period, review briefly the material covered in the lesson.

EVALUATION

Students are evaluated to ascertain whether they can demonstrate minimum competency with the lesson objectives in examination 1.

HOMEWORK

I. Complete work on the problems started above

II. Do the reading outlined in the course syllabus for lesson 7.

LESSON 6

STUDENT HANDOUT 6-1

Oxidation Numbers and Bond Polarities

The following exercise deals with; (1) the prediction of molucular formulae based upon common oxidation numbers and; (2) the prediction of bond polarity and the ordering of compounds based upon electronegativities.

1. Using the periodic table and your understanding of oxidation numbers, predict a molecular formula for each of the following pairs of atoms.

   Li & S

   Ca & O

   C & H

   Mg & F

   Aℓ & O

   Fe & O

2. Using the table of electronegativities in your text book and your understanding of electronegativities; (1) predict the degree of polarity of each molecule; (2) place them, relative to one another on the continuum; (3) identify the molecule which is likely to form with greatest ease.

   HCℓ

   H_2

   O_2

   KBr

   Aℓ As

   MgO

   HF

   Na F

   KF

   CsF.

LESSON 6

STUDENT ASSIGNMENT SHEET

For the Relationship of Periodic Classification to the Properties of the Elements

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 7. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 6 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 8.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
    - A. p. 45, 160, 133 (ionization energy)
    - B. Chapter 6, Section 6.8
    - C. Chapter 21, Section 21.1
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science, Chapters 5 and 21

PROBLEMS

1. Chapter 5; 5.15; 5.32; 5.35; 5.57
2. Chapter 21; 21.1; 21.14

TITLE General Concepts of Chemical Bonding II

LESSON NUMBER 7

TIME 50 minutes

DATE PREPARED February 15, 1981

BEHAVIORAL OBJECTIVES

As a result of participation in this lecture and completion of the associated activities, students should be able to:

I Identify principle and ℓ quantum numbers of bonding electrons involved in a bonding relationship

II Relate bonding electrons to bond types

III Draw sigma and pi bonds

I CONDITIONS

The conditions under which students must demonstrate minimum competency with the objectives of the lesson are:

A. using class handouts
B. using class notes
C. using the class text
D. by the end of the period
E. after having solved the problems listed below for this lesson

II CRITERIA

The level reflecting a demonstration of minimum competency with the objectives of this lesson is:

A. Students must answer correctly 90% of all objective related questions when exposed to them in a test situation.

REFERENCES

I INSTRUCTOR

A. Brown & LeMay
    1. Chapter 8, sections 8.3, 8.4, 8.5, 8.6

B. Solutions to Exercises in Brown & LeMay, problems 8.19, 8.38

II STUDENTS

A. Same as "A" above

B. Hill, chapter 8

MATERIAL/ EQUIPMENT

I INSTRUCTOR

A. chalk
B. chalkboard
C. periodic table
D. one each of the following handout for each student:
    1. student handout 7-1, Quantum Numbers in Chemical Bonding and Sigma and Pi Bonds

II STUDENT

A. paper
B. both "C" and "D" above

INTRODUCTION

This lesson is a continuation of lesson 6. Our discussion includes the relationship between quantum numbers, the octet rule and chemical bonding. In addition, we shall consider the formation of Sigma and pi bonds.

PROCEDURE

I LECTURE

A. Periodic table and oxidation numbers:

1\.

Family	Common Oxidation Number	Sample Elements
1A	+1	Na
7A	-1	Cl
2A	+2	Mg
6A	-2	O
7A	-1	F
2A	+2	Mg

2. Compound formation

a. $Na \longrightarrow Na^+ + 1e^-$
$1e^- + Cl \longrightarrow Cl^-$
$Na^+ + Cl^- \longrightarrow NaCl$

$\overset{+1}{Na}\ \ \overset{-1}{Cl}$ $(+1+{}^-1=0)$

b. $Mg \longrightarrow Mg^{2+} + 2e^-$
$1e^- + F \longrightarrow F$
$2e^- + 2F \longrightarrow 2F$
$Mg^{2+} + 2e^- + 2F \longrightarrow MgF_2$

$\overset{2^+}{Mg}\ \overset{(2x^-1)}{F_2}$ $(+2+{}^-2=0)$

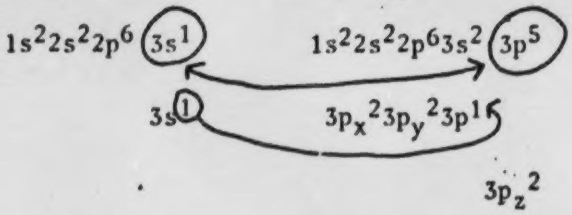
B. Periodic table and outer orbital formulae:

Family	Outer Orbital Formula	Sample Elements
1A	ns^1	Na
7A	ns^2np^5	Cl

1. $ns^2np^5 + 1e^- = ns^2np^6$
   - a. electron total 2 + 6 = 8
   - b. $8e^-$ = octet rule
   - c. $\underset{3s^1}{Na} \longrightarrow \underset{2s^22p^6\ (8e^-)}{Na^+} + 1e^-$
   - d, $1e^- + \underset{3s^23p^5}{Cl} \longrightarrow \underset{3s^23p^6 (8e^-)}{Cl^-}$

C. Bonding

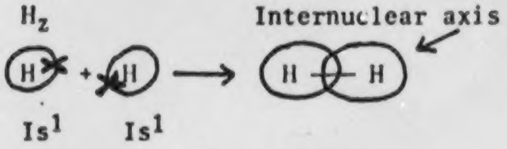
1. Na + Cl



$1s^22s^22p^6 3s^1$
$1s^22s^22p^63s^2 3p^5$
$3s^1$
$3p_x{}^2 3p_y{}^2 3p^1$
$3p_z{}^2$

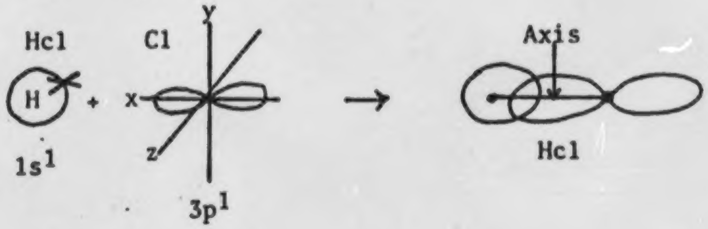
2. Sigma Bonds - bonds where the portion of the electron cloud created by a bonding pair of electrons surrounds the internuclear axis

a. H_2



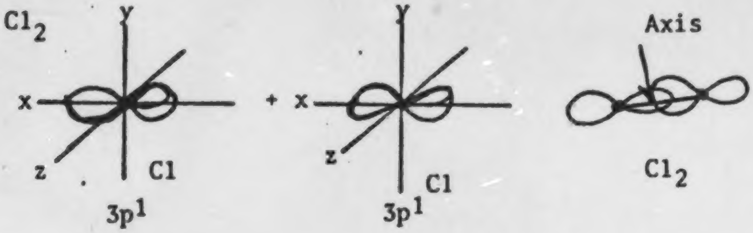
Internuclear axis
H
H
H
H
$1s^1$
$1s^1$

b. HCl



Cl
y
Axis
H
x
z
$1s^1$
$3p^1$
Hcl

c. Cl_2



y
y
Axis
x
x
z
z
Cl
Cl
$3p^1$
$3p^1$
Cl_2

3. Pi Bonds - bonds where the electron density created by the bonding pair of electrons does not surround the internuclear axis.

Axis

Axis

END VIEW

4. Bond Hybridization

BOND TYPE	CONSTITUENT	BOND ANGLE	GEOMETRY	EXAMPLE	NUMBER OF ORBITALS FORMED
SP	1S & 1P	180°	linear	$BeCl_2$	2
SP^2	1S & 2P	120°	tri. planar	BCl_3	3
SP^3	1S & 3P	109.5°	tetra hedral	CH_4	4

II ACTIVITY

A. Divide students in groups of 4.

B. Have each group of students do the activity outlined in student handout 7-1.

C. Have each group of students begin work on text book problems 8.19 and 8.38.

APPLICATION Students apply background gained here to future work in the course.

SUMMARY At the end of the session, review briefly the material covered in the lecture section of the session.

EVALUATION Students are evaluated to determine whether they can demonstrate minimum competency with the objectives of this lesson by reviewing the results of exam 1.

HOMEWORK

I. Complete work on problems started above.

II. Do the reading outlined on the course syllabus for lecture 8.

Lesson 7

STUDENT HANDOUT 7-1

Quantum Numbers in Chemical Bonding and Sigma and Pi Bonds

DIRECTIONS:
Based upon the material presented in this lesson, class handouts and information presented in relevant sections of your class text determine for each compound presented below:

(1) the principle quantum number and suborbital (s, p, d etc.) type in which each bonding electron resides prior to bond formation; (2) the type of bonding orbital formed; (3) Lewis structure for each compound.

Rb_2O CsF C_2H_4 (C=C)

CaS $SrC\ell_2$ $A\ell_2O_3$

SiO_2 $A\ell F_3$ FeO

LESSON 7

STUDENT ASSIGNMENT SHEET

For General Concepts of Chemical Bonding I

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 8. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 7 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 9.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
    - A. Chapter 2, Section 2.8
    - B. Chapter 7, Section 7.1; 7.2; 7.4; 7.5; 7.9; 7.10
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science, Chapter 7

PROBLEMS

1. Chapter 2; 2.28; 2.32; 2.52a; 2.57
2. Chapter 7; 7.1; 7.2; 7.4; 7.7; 7.10

TITLE Examination Review

LESSON NUMBER 8

TIME 50 minutes

DATE PREPARED February 14, 1981

BEHAVIORAL OBJECTIVES

As a result of completing the activity associated with this lesson students should be able to:

I. Answer correctly all of the questions on Exam 1.

I CONDITIONS

The conditions under which students must demonstrate minimum competency with the objective are:

A. Using class notes
B. Using class handouts
C. Using class text
D. Using laboratory manual
E. By the end of the class period

II CRITERIA

Minimum competence with the objective stated above is demonstrated when students answer correctly all questions on Exam 1.

REFERENCES

I INSTRUCTOR None

II STUDENTS

A. Review all reading in Brown & LeMay required to date
B. Problems solved to date

MATERIALS/EQUIPMENT

I INSTRUCTOR

A. One copy of student handout 8-1, Examination I Review for each student

II STUDENTS

A. Paper

INTRODUCTION

This session is intended to be a review. The questions included in Student Handout 8-1 wire designed to help students meet and demonstrate minimum competency with the lectures conducted since the beginning of the course.

PROCEDURE

I Activity

A. Divide students into groups of four
B. Have students work cooperatively to solve problems on Exam 1.

SUMMARY

At the end of the class, review the relationship between the questions posed in Student Handout 8-1 and the objectives of the previously presented lessons.

EVALUATION

Students' competence with the objective of this lesson is evaluated by examining the results of Exam 1.

HOMEWORK

Continue to prepare for Exam 1.

STUDENT HANDOUT 8-1
Examination I Review
Lesson 8

DIRECTIONS: The following group of problems were designed around the behavioral objectives dealt with during the first third of the course. They are intended to help you review the course material covered to date. You should set each problem up and solve it exactly as if this were a real examination.

1. Using the concept of dimensional analysis convert 1 m^3 to cubic millimeters.

2. Define chemical properties of matter.

3. Define physical properties of matter.

4. A sample of some unknown has a density of 1.7g/cm^3 and a volume of 10 ml. What is the mass of the sample?

5. What is the temperature in °F of a substance if it was measured at 202°K.

6. Outline and define the divisions matter from heterogeneous and homogeneous to elements.

7. If M C
E
A Q

Draw a symbol and set of numbers which reflects an oxygen atom which has gained two electrons.

8. Determine the number of neutrons possessed by the element in the fifth period and family 7A.

9. How many electrons can possibly exist at n=5?

10. Write the orbital formula for elements 53.

11. Draw, side-by-side a diagram of $4p_x$, $4p_y$, $4p_z$ and $5p_x$, $5p_y$, $5p_z$ suborbitals.

12. Which of the following suborbitals fills first, 5s or 4d and which has the greatest amount of energy?

it
13. Define the oclet rule & how/relates it to common oxidation states.

required
14. Why do group 3A elements show a large increase in/ionization energy between required
the third and forth ionization energies.

15. Predict the probable molecular formulae between family 1A elements and group 5A elements.

16. What type of bonding orbital exists in the compound HCℓ?

17. What is the percent of sodium in a compound if a 7.69g sample is found to contain 2.28g of Na?

LESSON 8

STUDENT ASSIGNMENT SHEET

For General Concepts of Chemical Bonding II

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 9. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 8 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 10.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
    - A. Chapter 8, Sections 8.3; 8.4; 8.5; 8.6
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science, Chapter 8.

PROBLEMS 1. Chapter 8, 8.19; 8.38

TITLE	Examination 1
LESSON NUMBER	9
TIME	50 minutes
DATE PREPARED	February 16, 1981
BEHAVIORAL OBJECTIVE	As a result of taking this examination students should be able to: I Identify areas of conceptual weakness with course material covered heretofore.
I CONDITIONS	A. Following correction of the exam
II CRITERIA	None
REFERENCES	
I INSTRUCTOR	None
II STUDENTS	A. corrected exam
MATERIALS/EQUIPMENT	
I INSTRUCTOR	A. One copy of Exam 1 for each student B. One periodic table for each student
II STUDENTS	A. "A" and "B" above
INTRODUCTION	Refer students to the instructions on page 1 of the exam.
PROCEDURE	Administer the exam
APPLICATION	Students use information gained from the exam to correct any misunderstandings they may have over the material covered thus far in the course.
SUMMARY	None
EVALUATION	None
HOMEWORK	Students do reading outlined on the course syllabus for lesson 10.

Spring 1981

Ch-1 INTRODUCTORY COLLEGE CHEMISTRY

Examination 1

Name:____ Date:_____________

Directions: Answer each of the questions in the space provided. If there is more than one step involved in the solution, show each step and number the steps successively. It is a good idea to answer the questions you are sure about first and then ponder the rest. You have 50 minutes to complete the exam. You may use your text books, class notes and laboratory manuals as you answer the questions.

1. When full, a ships fuel oil day tank holds 100 cubic yards at 20^{o} X C. If 1m equals 1.094 yards, using dimensional analysis, show how this volume of fuel can be converted to cubic meters.

2. What is the volume of a liquid that weights 25.2 g and has a density of 1.5 g/ml.

answer________________

3. Which of the following can be classified as physical properties?

a) melting point
b) temperature
c) color
d) volume
e) mass
f) density

answer________________

4. Six elements respetively have the following numbers of protons 12, 5, 7, 2, 16, 9. Based upon these data, provide the following information.

	Atomic Number	Symbol	Family	Period	No. of Neutrons	No. of Electrons
1.						
2.						
3.						
4.						
5.						
6.						

5. Provide answers to the questions associated with the following symbol:

$$^{118}_{50}\mathrm{Sn}^{4+}$$

a. Name the element ______________________________
No. of neutrons ______________________________
Common oxidation state ______________________________
Principle quantum number of the valence electrons ______________

6. Write the complete orbital formula for antimony. The formula should show proper filling order.

7. Draw a picture reflecting the difference you would expect to find between the 2px and the 3px suborbitals.

8. Provide answers to the following questions associated with the following:

$15^2\ 2s^2\ 2p^6\ 3s^2$

a. Name of the element ______________________
b. Suborbital farthest from the nucleus ______________________

9. Why is the fourth ionization energy of Boron so much greater than the third?

10. Predict molecular formulae between the following pairs of atoms using the common oxidation number when no other is indicated.

Cu (+ 2 & +3) & As   Na & S

11. Six elements have the following electronegativities Na.93, Cℓ 3.2, Li 1.0, Br 3.0, K.82, F 4.0. Which of the following molecules is most polar and which is least polar?

Li F, K Br, Na Cℓ

12. Whey do we calibrate thermometers and pipets?

13. Why do we never weigh hot objects?

14. What is the percentage of SiO_2 in a 8.69g sample mixture if 4.76g of SiO_2 has been recovered (4.76 is assumed to be all the SiO_2 held by the sample).

LESSON 9

STUDENT ASSIGNMENT SHEET

For Examination I Review

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 10. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 9 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 11.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
   - A. All reading required to date

PROBLEMS 1. Review problems solved to date

TITLE Conservation of Mass in Chemical Systems

LESSON NUMBER 10

TIME 50 minutes

DATE PREPARED February 16, 1981

BEHAVIORAL OBJECTIVES As a result of participation in this lecture and completion of the associated activities students should be able to:

I Define molecular empirical or simplest and structural formulae

II Define chemical equation

III Relate the Avagadro number to gram molecular weights and the periodic table

IV Compute molecular mass and composition from molecular formulae

I CONDITIONS The conditions under which students must demonstrate minimum competencey with the objectives are:

A. Using class notes
B. Using class handouts
C. Using the class text
D. By the end of the class period
E. After having solved the problems listed below for this lesson

II CRITERIA The level reflecting a demonstration of minimum competence with the objectives of this lesson is:

A. Students must answer correctly 90% of all objective related questions when exposed to them in a test situation.

REFERENCES

I INSTRUCTOR

A.
Brown & LeMay 1. Chapter 2, section 2.8

2. Chapter 3, section 3.4, 3.6, 3.7, 3.8, 3.9, 3.1, 3.2, 3.3

3. Kolb, D. the mole. Journal of Chemical Education 728-732, 55(II), 1978

B.
Solutions to Exercises in Brown and Lemay, problems 3.10, 3.12, 3.15, 3.22, 3.47.

II STUDENTS

A. Same as "A" above

B. Hill, Chapter 3.

MATERIALS/EQUIPMENT

I INSTRUCTOR

A. Chalk
B. Chalkboard
C. Periodic table
D. One copy of student handout 10-1, the Mole for each student
E. 12 each of the following items
    1. marbles
    2. toothpicks
    3. 10 penny nails
    4. 1/4 x 20 nuts
    5. paper clips
    6. number 10 flat washers
    7. 1/2" flat washers
    8. 1/4" x 20 x 1/2" hexagon head bolts
F. Eight containers, one for each of the groups of objects listed above in E.
G. Eight triple arm balances

II STUDENTS

A. Paper

INTRODUCTION

An understanding of conservation of mass in chemical systems relies upon an anticedent understanding of the molar concept. In this lecture, we begin a discussion of that concept and related definitions. During the session we will switch from lecture to student activity to lecture.

PROCEDURE

LECTURE

A. Define the following terms

1. Molecular formula-formulae which indicate the actual numbers and types of atoms present in a molecule
    a. Hcl
    b. $NaOH$
    c. NH_3

2. Empirical or simplest formula-formulae which give only the relative numbers of atoms not the actual number:

    a. HO H_2O_2 b. CH_2 C_2H_4

3. Structural formula-formulae showing the relative arrangement of atoms

4. Chemical equation-symbols and formulas that combine to make a statement:

    a. $C + O_2 \rightarrow CO_2$
    b. $NaOH + HCl \longrightarrow H_2O + NaCl$

II ACTIVITY

A. Divide students in groups of four

B. Have each group conduct the activity outlined in Student Handout 10-1.

1. Allow 15 minutes for the investigative portion of the activity
2. Appoint one person in each group chairman. Chairman will report the results of the activity to the entire class.

C. Class discussion of the activity.

1. Each group reports.
2. Relate findings to the periodic table
3. Define the Avagadro number
4. Define gram molecular weight.

III LECTURE

A. Percent composition from formulae

1. Example:

$$1H + 1Cl \rightarrow 1HCl$$
$$1 \text{ mole } H + 1 \text{ mole } Cl \longrightarrow 1 \text{ mole } HCl$$
$$1gH + 35.5gCl \rightarrow 36.5gHCl$$

? % composition of H

$$1gH \times \left(\frac{100\%}{36.5gH}\right) = \% \text{ composition}$$

APPLICATION

Students apply skills and understanding gained in this lesson to work in lesson 11.

SUMMARY

At the end of the class, briefly summarize the main points of the lesson.

EVALUATION

Students are evaluated to determine whether they meet minimum competency with the objectives of this lesson by reviewing the results of exam 2.

HOMEWORK

I. Have students solve textbook problems 3.10, 3.12, 3.15, 3.22, 3.47.

II. Have students do the reading listed in the course syllabus for lesson 11.

Lesson 10

Student Handout 10-1
Molar Concept

DIRECTIONS:

I. You have a group of 12 objects and a balance in front of you. Using your intuition as well as the equipment provided evaluate as many properties of these objects as is possible. As you consider the objects, look for and evaluate:

A. The variables common to each object.

B. The variables which may be held constant across the group while other variables are allowed to change.

II. Answer the following questions.

A. What characteristics of the objects are different.

B. What characteristics of the objects are the same.

C. What mathematical, measuring and graphing tools would you use to investigate the relationship between mass, number and size.

D. What would be the result of multiplying the number of items by

$.5016 \times 10^{23}$?

E. What is the relationship between the answer to question D and the chemist dozen (see chart below)?

ITEM	KIND OR SET	NUMBER IN SET
Sox, dice	pair	2
Eggs, oranges	dozen	12
Bottles, cans	case	24
Brushes, pencils	gross	144
Sheets of paper	ream	500
Atoms, molecules	mole	$6.02\text{X}10^{23}$

III. Answer the following questions:

A. List the mass of all second period elements. Are their differences in these masses and if so, what might lead to these differences?

B. How many atoms would you hypothesize to be in the mass in grams of each element in period 2 and what led you to that conclusion?

C. In answering B, what variables was it necessary to consider, which ones remained constant and which varied acorss the period. What are the reasons for the variation.

TITLE Conservation of Mass in Chemical Systems II

LESSON NUMBER 11

TIME 50 minutes

DATE PREPARED February 27, 1981

BEHAVIORAL OBJECTIVES As a result of attending this lecture and completing the associated activities, students should be able to:

I Describe the meaning of conservation of mass in terms of:

A. Molecular coefficients
B. Moles of molecules
C. Mass in grams

I CONDITIONS The conditions under which students must demonstrate minimum competency with the objectives are:

A. Using class notes
B. Using class handouts
C. Using the class text
D. By the time Examination Two is administered

II CRITERIA The level reflecting a demonstration of minimum competency with the objectives is:

A. Students must answer correctly 90% of all questions related to the objectives of this lesson.

REFERENCES

I INSTRUCTOR

A. Brown & LeMay - same as reading for lesson 10.

B. Student Guide to Brown & LeMay, Chapter 3.

C. Answers to Problems in Brown and LeMay, problems 3.2, 3.3, 3.7a, c, e, 3.4.

D. Kolb, D. The Mole *Journal of Chemical Education*, 728-732, 55(11), 1978.

II STUDENTS A. Same as A & B above

MATERIAL/EQUIPMENT

I INSTRUCTOR

A. Chalkboard
B. Chalk
C. Periodic table
D. Student handouts - one copy of the following handout for each student:
    1. Student Handout 11-1, Conservation of Mass in Chemical Systems
    2. Student Handout 11-2, The Mole and Avagadro's Number

II STUDENTS

A. Paper
B. Both C & D above

INTRODUCTION

In chemical systems, elements combine in definite proportions to one another as they form compounds. Their ratios by mass are constant and predictable. During the process called a chemical reaction mass is neither gained nor lost but conserved. We shall consider these tenants throughout the remainder of this lesson.

PROCEDURE

I. LECTURE

A. Law of Definite proportions - the ratios of combining elements by mass are constant.

a. NaCl (combination reaction)

$$A + B \longrightarrow AB$$
$$Na + Cl \longrightarrow NaCl$$

1 atom/1 atom/atom ratio 1:1

$\frac{23 \text{ mass units}}{23}$ $\frac{35.5 \text{ mass units}}{23}$ mass ration 1:1.54

22.9gNa = 1 mole Na
35.4gCl = 1 mole Cl

Family 1a $Na \longrightarrow Na^+ + 1e^-$
Family 7a $1e^- + Cl \longrightarrow Cl^-$

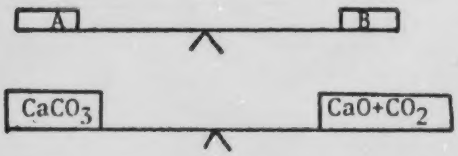
B. Law of Conservation of Mass - mass is neither gained nor lost in a chemical reaction

1. Equation (decomposition)

$$AB \longrightarrow A + B$$

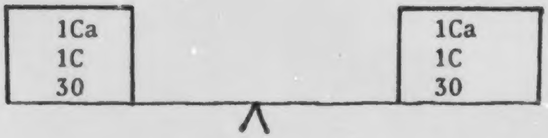
$$\underbrace{CaCO_3(s)}_{\text{Reactant}} \xrightarrow[\text{or changes to}]{\text{yields}} \underbrace{CaO(s) + CO_2(g)}_{\text{Products}}$$

a. Mass of Side A = Mass of Side B

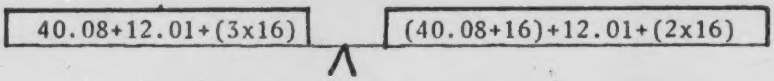


Element	Atomic Mass
Ca	40.08
C	12.01
0 x 3	48.0

Balance by atoms



Balance by mass



2. The balanced equation and the mole

a. Example (Replacement or Substitution)

$$A \quad + \quad BC \longrightarrow AB \quad + \quad C$$

$$\underbrace{25\underbrace{O_2}_{\text{1 mole}}}_{\text{25 moles}} + \underbrace{2\underbrace{C_8H_{18}}_{\text{1 mole}}}_{\text{2 moles}} \qquad\qquad \underbrace{16\underbrace{CO_2}_{\text{1 mole}}}_{\text{16 moles}} + \underbrace{18\underbrace{H_2O}_{\text{1 mole}}}_{\text{18 moles}}$$

$$\begin{array}{cccc} \downarrow & \downarrow & \downarrow & \downarrow \\ 25(6.02\text{x}10^{23}) & 2(6.02\text{x}10^{23}) & 16(6.02\text{x}10^{23}) & 18(6.02\text{x}10^{23}) \\ \text{(molecules)} & \text{(molecules)} & \text{(molecules)} & \text{(molecules)} \end{array}$$

b. Reactants

(1) $25O_2$

$$25\left[\begin{array}{l} 2\ (6.02 \text{ x } 10^{23}) \\ \quad \text{O atoms} \end{array}\right]$$

$50(6.02 \text{ x } 10^{23})$ O atoms

$30 \text{ x } 10^{23}$ O atoms

(2) $2C_8H_{18}$

$$2\left[\begin{array}{lcl} 8(6.02 \text{ x } 10^{23}) & + & 18(6.02 \text{ x } 10^{23}) \\ \quad \text{C atoms} & & \quad \text{H atoms} \end{array}\right]$$

$$\begin{array}{lcl} 16(6.02 \text{ x } 10^{23}) & + & 36(6.02 \text{ x } 10^{23}) \\ \quad \text{C atoms} & & \quad \text{H atoms} \end{array}$$

$96.32 \text{ x } 10^{23}$ C atoms

$216.72 \text{ x } 10^{23}$ H atoms

c. Products

(1) $16CO_2$

$$16\left[\underset{\text{C atoms}}{(6.02 \times 10^{23})} + 2\underset{\text{O atoms}}{(6.02 \text{ X } 10^{23})}\right]$$

$$16\underset{\text{C atoms}}{(6.02 \times 10^{23})} \quad + \quad 32\underset{\text{O atoms}}{(6.02 \times 10^{23})}$$

96.32×10^{23} C atoms
192.64×10^{23} O atoms

(2) $18H_2O$

$$18\left[2\underset{\text{H atoms}}{(6.02 \times 10^{23})} \quad + \quad \underset{\text{O atoms}}{(6.02 \times 10^{23})}\right]$$

$$36\underset{\text{H atoms}}{(6.02 \times 10^{23})} \quad + \quad 18\underset{\text{O atoms}}{(6.02 \times 10^{23})}$$

216.72×10^{23} H atoms
108.36×10^{23} O atoms

d. The balance

(1)

Reactants	Products
301.0×10^{23} O atoms	192.64×10^{23} O atoms
216.72×10^{23} H atoms	$\underline{108.36\times10^{23}}$ O atoms
96.32×10^{23} C atoms	301.00×10^{23} O atoms
	216.72×10^{23} H atoms
	96.32×10^{23} C atoms

II ACTIVITY

A. Divide students into groups of four

B. Have each group complete the activity outlined on Student Handout 11-1, Conservation of Mass in Chemical Systems

APPLICATION

I. Students apply the background gained in this lecture to:
   A. Lesson 12
   B. Examination 2
   C. Future work in chemistry

SUMMARY

At the end of the class period, review briefly the molar concept using Student Handout 11-2, The Mole and Avagadro's Number

EVALUATION

Students are evaluated in examination two to determine whether they can demonstrate minimum competency with the objectives of this lesson.

HOMEWORK

I Complete the work started above in activity section

II Solve problems: 3.2, 3.3, 3.7a, c, e, 3.4

III Have students do the required reading for lesson 12 (on course syllabus)

LESSON #11

STUDENT HANDOUT 11-1

Conservation of Mass in Chemical Systems

DIRECTIONS:

Based upon material presented in this lesson, class handouts, information presented in relevant sections of your class text determine the following. With each of the following chemical equations, ascertain whether they are balanced by moles, atoms and mass in grams. In addition to the information sources listed above, you will need a periodic table.

1. $3s + 4HNO_3 \rightarrow 3SO_2 + 4NO_2 + 2H_2O$
2. $10HNO_3 + I_2 \rightarrow 2HIO_3 + 10NO_2 + 4H_2O$
3. $KMnO_4 + Hc\ell \rightarrow MnC\ell_2 + C\ell_2 + KC\ell + H_2O$
4. $NH_4OH + HC\ell \rightarrow NH_4C\ell + H_2O$

LESSON #11

STUDENT HANDOUT 11-2

The Mole and Avogadro's Number

Avogadro's number is 6.02×10^{23}. It is an unchanging number just as the dozen is unchanging. When atomic masses of the elements are taken in grams each of those masses contains the avogadro number of atoms. When the formula weights of compounds (in atomic mass units) are taken in grams, those masses contain the avogadro number of molecules.

Element	Atomic Number	Atomic Weight	Grams/ Mole	Atoms/ Mole
H	1	1	1	6.02×10^{23}
He	2	4	4	6.02×10^{23}
Li	3	7	7	6.02×10^{23}
BE	4	9	9	6.02×10^{23}
B	5	11	11	6.02×10^{23}
C	6	12	12	6.02×10^{23}
N	7	14	14	6.02×10^{23}
O	8	16	16	6.02×10^{23}
F	9	19	19	6.02×10^{23}
Ne	10	20	20	6.02×10^{23}
Na	11	23	23	6.02×10^{23}
Mg	12	24	24	6.02×10^{23}

Compound	Molecular Weight	Molecules/Mole
H_2	2g	6.02×10^{23}
H_2O	18g	6.02×10^{23}
$NaC\ell$	58.43g	6.02×10^{23}
$HC\ell$	36.45g	6.02×10^{23}
NaOH	39.75g	6.02×10^{23}

LESSON II

STUDENT ASSIGNMENT SHEET

For Conservation of Mass in Chemical Systems I

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 12. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 11 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 13.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
   - A. Chapter 2; 2.8
   - B. Chapter 3, Sections 3.4; 3.6; 3.7; 3.8; 3.9; 3.2; 3.2; 3.3
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science, Chapter 3

PROBLEMS 1. Chapter 3; 3.10; 3.12; 3.15; 3.22; 3.47

TITLE Conservation of Mass in Chemical Systems III

LESSON NUMBER 12

TIME 50 minutes

DATE PREPARED March 2, 1981

BEHAVIORAL OBJECTIVES As a result of attending this lecture and completing the associated activities students should be able to:

I Relate the following concepts to chemical equations:
- A. Molarity
- B. Normality
- C. Equivalent Weights
- D. Molality
- E. Weight Percent
- F. Limiting Reagent

I CONDITIONS The conditions under which students must demonstrate minimum competency with the objectives are:

- A. Using class notes
- B. Using class handouts
- C. Using the class text
- D. By the time examination two is administered

II CRITERIA The level reflecting a demonstration of minimum competency with the objectives is:

- A. Students must answer correctly 90% of all questions related to the objectives of this lesson.

REFERENCES

I INSTRUCTOR A. Brown & LeMay

1. Chapter 12, Section 12.1
2. Chapter 3, p 82 (limiting reagent)

B. Student Guide to Brown & LeMay, Chapters 3 and 12

C. Answers to Problems in Brown & LeMay, problems 12.1, 12.3, 3.11, 3.77

II STUDENTS A. Same as A & B above

MATERIALS/ EQUIPMENT

I INSTRUCTOR
- A. Chalk
- B. Chalkborad
- C. Periodic table

II STUDENTS

A. Paper
B. Periodic table

INTRODUCTION

Quantification in chemistry requires the use of a number of different measuring methods. We will consider these methods today and then consider how at least one (molarity) may be used to determine which reagent in a chemical reaction is the limiting reagent.

PROCEDURE

I. LECTURE

A. $\text{Molarity(M)} = \frac{\text{Moles of Solute}}{\text{liters of solution}}$

1. 39.97g NaOH in 1l of H_2O=1 molar solution

2. 98gH_2SO_4 in 1 l of H_2O= 1 molar solution

B. $\text{Normality(N)} = \frac{\text{Equivalent weights}}{\text{liters of solution}}$

1. $\text{Equivalent Weight} = \frac{\text{Molar Weight}}{N}$

2. N = number of e^- gained or lost by a substance

3. N = number of H^+ provided by a molecule

4. N = number of OH^- provided by a formula unit of base

5. One Equivalent Weight = Mass of substance which will produce the Avagadro number of cahrges (6.02×10^{23}).

C. $\text{Molality} = \frac{\text{Moles of Solute}}{\text{1000g of Pure Solvent}}$

1. pure solvent = a fixed mass in moles/Kg

2. not affected by temperature

D. Percent by Weight or Mass Percent =

$$\frac{\text{Mass of Solute}}{\text{Mass of Solute+Mass of solvent}} \times 100$$

E. Limiting Reagent = the reagent used up first which prevents the reaction from going to completion

1. Example HCl + NaOH

   a. $HCl + NaOH \rightarrow NaCl + H_2O$

   b. $HCl \rightarrow H^+ + Cl^-$

   c. $NaOH \rightarrow Na^+ + OH^-$

   d. 1 mole NaOH + 1 mole HCl   1 mole NaCl + 1 mole H_2O

   e. .5 mole NaOH+1mole   ½ mole NaCl+½ mole H_2O
   ½ mole HCl left in the reaction vessel

II ACTIVITY

A. Divide the class into groups of four

B. Have students answer the questions listed on Student Handout 11-1.

C. If time permits, have students begin homework problems 12.1, 12.3, 3.11 and 3.77

APPLICATION

I Students apply the background gained in this lesson to:

A. Lesson 13
B. Examination 2
C. Future work in chemistry

SUMMARY

At the end of the class period, review briefly the definitions of molarity, normality and equivalent weights.

EVALUATION

Students are evaluated to see if they can demonstrate minimum competency with the objectives of this lesson by reviewing the results of Examination Two.

HOMEWORK

I. Complete the work started in the activity session

II. Solve problems 12.1, 12.3, 3.11, 3.77

III Do the required reading for lesson 13 (on course syllabus)

LESSON 12

STUDENT HANDOUT 12-1

Quantitive Measurements
Used In Chemistry

DIRECTIONS:

Study the examples provided below and then answer the questions following the examples. In answering the questions, you should use your class notes, text, handouts and so on.

Examples:

1. Molarity, normality & equivalent weights

   a) $Hc\ell$

      (1) $Hc\ell \rightarrow H^+ + C\ell^-$

      (2) /mole $Hc\ell$ → /mole H^+ +/mole $C\ell$=

      (3) /mole/$\ell Hc\ell$ = 1M solution

      (4) mole/$\ell Hc\ell$ = 1N solution because 1 mole of H+ are produced

      (5) 1 equivalent weight $Hc\ell$ = 36.4g $Hc\ell$ = 1 mole $Hc\ell$

   b) NaOH

      (1) $NaOH \rightarrow Na^+ + OH^-$

      (2) 1 mole NaOH → 1 mole Na^+ + 1 mole OH^-

      (3) 1 mole/ℓ NaOH = 1M solution

      (4) 1 mole/ℓ NaOH = 1N solution because 1 mole of OH^- are produced

      (5) 1 equivalent weight NaOH = 39.8g NaOH = 1 mole NaOH

   c) H_2SO_4

      (1) $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$

      (2) 1 mole H_2SO_4 →2 moles H^+ + 1 mole SO_4^{2-}

      (3) 1 mole/ℓ H_2SO_4 = 1M solution

      (4) 1 mole/ℓ H_2SO_4 = 2N solution because 2 moles of H^+ are produced

      (5) 1 equivalent weight of H_2SO_4 = 98g/2

(6) $.5M/\ell H_2SO_4 = 1N$ solution

2. Limiting Reagents

a) $NH_4OH + HC\ell \rightarrow NH_4C\ell + H_2O$

(1) 17g NH_4OH+ 36.5g $HC\ell$

(2) NH_4OH is the limiting reagent

Questions:

From the information given in each equation, as applicable, determine; (1) the normality of each reactant; (2) molarity of each species, both reactants and products; (3) the equivalent weights of each reactant.

1. 49g NaOH + 35g$Hc\ell \rightarrow NaC\ell + H_2O$

2. $2\ NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$

3. $A\ell(OH)_2 + Hc\ell\ A\ell C\ell_2 + H_2O$

4. $Zn\ (OH)_2 + 4Hc\ell \rightarrow H_2O + Zn\ C\ell_2$.

II. Answer the following questions.

1. How much $Hc\ell$ must be reacted with 80g of NaOH if the reaction goes to completion?

2. If 148.9g of $KC\ell$ are dissolved in 2 liters of water, what is the molarity of the solution?

3. In question 2, what is the mass an equivalent weight?

4. 98g H_2SO_4 are dissolved in .25ℓ of solution. What is the normality and what is the molarity of the solution?

LESSON 12

STUDENT ASSIGNMENT SHEET

For Conservation of Mass in Chemical Systems II

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 13. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 12 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 14.

PROBLEMS 1. Chapter 3; 3.2; 3.3; 3.7a, c, e; 3.4

TITLE Problems and the Conservation of Mass

LESSON NUMBER 13

TIME 50 minutes

DATE PREPARED March 2, 1981

BEHAVIORAL OBJECTIVES As a result of attending this lecture and completing the associated activities students should be able to:

I Solve reactions by mass problems
II Solve atoms to mass problems
III Solve mass to moles problems
IV Solve atoms of A to grams of B problems

I CONDITIONS The conditions under which students must demonstrate minimum competency with the objectives are:

A. Using class notes
B. Using class handouts
C. Using the class text
D. By the time Examination Two is administered

II CRITERIA The level reflecting a demonstration of minimum competency with the objectives is:

A. Students must answer correctly 90% of all questions related to the objectives of this lesson

REFERENCES

I INSTRUCTOR A. Brown & LeMay

1. Sample exercises 3.5, 3.6, 3.7, 3.8, 3.9, 3.10, 3.11, pp 73-78

B. Students Guide to Brown & LeMay, Chapter 3.

C. Answers to Problems in Brown & Lemay, problems 3.21, 3.28, 3.30.

II STUDENTS A. Both A and B above

MATERIALS/EQUIPMENT

I INSTRUCTOR
A. Chalk
B. Chalkboard
C. Periodic table
D. One copy for each student of Student Handout 13-1, Conservation of Mass Problems

II STUDENTS
A. Paper
B. Periodic Table

INTRODUCTION Today we begin applying the concepts of the mole, conservation of mass and the like to real world situations

PROCEDURE **I LECTURE**

A. Reactions by mass
  1. Mass is conserved
  2. All reactants are consumed-reaction goes to completion
  3. Products are neutral
  4. Example:

$36.54g\ Hcl + 39.97g\ NaOH \longrightarrow$

$17.99g\ H_2O + 58.43g\ NaCl$

  a. An acid-base neutralization
  b. Same reaction by molecules

$6.02 \text{ x } 10^{23}Hcl + 6.02 \text{ x } 10^{23}NaOH \longrightarrow$

$6.02 \text{ x } 10^{23}H_2O + 6.02 \text{ x } 10^{23}NaCl$

B. Atoms to mass problems
  1. Example - a single razor blade contains $8.4 \text{ x } 10^{21}$ atoms; 57% of the atoms are Fe, 14 P/O are Cr and 29% are C. What is the mass of carbon contained in the blade?

  2. Steps in solution

  a. 57% + 29% + 14% = 100% (total mass)

  b. $(8.4x10^{21}\text{atoms})(\frac{29c\ \text{atoms}}{100\ \text{atoms}})(\frac{1\ \text{mole C atoms}}{6.02x10^{23}\text{C atoms}})$

$(\frac{12g\ C}{1\ \text{mole } C}) = 4.9 \text{ x } 10^{-2}{}_{g}$ Carbon

C. Mass to Moles
  1. Example - the mass of an unknown atom is $2.40 \text{ x } 10^{-22}$g. What fraction of a mole of the unknown is represented by a sample weighing 10.5g?

  2. Steps in solution

  a. $(10.5gX)(\frac{1\ \text{atom } x}{2.40x10^{-22}{}_{g}x})(\frac{1\ \text{mole } x}{6.02x10^{23}\text{atoms}x}) = 7.27$ moles X

D. Atoms of A to Grams of B
  1. Example - there are $8.03x10^{23}$ atoms of N in a sample of $(NH_4)_2Cr_2O_2$. How many grams of H are in the sample?

  2. Steps in solution

  a. $(8.03x10^{23}\text{atoms N})(\frac{1\ \text{mole N}}{6.02x10^{23}\text{atoms N}})$

$(\frac{8\ \text{moles H}}{2\ \text{moles N}})(\frac{1g\ \ H}{1\ \text{mole H}}) = 5.39g$ H present

II ACTIVITY

A. Divide the class into groups of four students

B. Have each group set up and solve the problems listed on student Handout 13-1.

C. If time remains have each group begin to solve homework problems 3.21, 3.28, 3.30.

APPLICATION

I Students apply the skills gained here to:

A. Lesson 14
B. Future work in chemistry

SUMMARY

At the end of the session, review briefly the importance of dimensional analysis in the solution of such problems.

EVALUATION

Students are evaluated to determine whether they can demonstrate minimum competency with the objectives by reviewing the results of Examination 2.

HOMEWORK

I Complete the activity started above.

II Solve problems 3.21, 3.28, 3.30

III Do the required reading for Lesson 14 (on course syllabus).

LESSON #13

STUDENT HANDOUT 13-1

Conservation of Mass Problems

DIRECTIONS:

Using the same logic system dimensional analysis and the like as just used in the lecture, set up and solve the following problems. Solutions to these problems are found on the back of this sheet but before you look at them, try your hand at it.

1. A bottle weighing 6.00g contains an unknown number of carbon atoms. Another bottle contains an equal number of nickel atoms. What is the mass of the Ni atoms?
2. What mass of $KC\ell O_3$ contains 80g of oxygen?
3. If $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$, how many O molecules are required to react with 68g NH_3?
4. If $Ca_3P_2 + 6H_2O \rightarrow 3Ca(OH)_2 + 2PH_3$, how much PH_3 can be prepared from 2.0g Ca_3P_2?

LESSON #13

STUDENT HANDOUT 13-A

Solutions to the 13-1 Conservation of Mass Problems

1. $$(6.00\text{gC})\left(\frac{1\text{ Mole C}}{12\text{ grams C}}\right)\left(\frac{1\text{ mole Ni}}{1\text{ mole C}}\right)$$

$$\left(\frac{58.7\text{g Ni}}{1\text{ mole Ni}}\right) = 29.35\text{g Ni}$$

2. $$(80.0\text{g}0)\left(\frac{1\text{ mole }0}{16\text{g }0}\right)\left(\frac{1\text{ mole KC}\ell 0}{3\text{ moles }0}\right)$$

$$\left(\frac{122.5\text{g KC}\ell 0_3}{1\text{ mole KC}\ell 0_3}\right) = 204.16\text{g KC}\ell 0_3$$

3. $$(68\text{g NH}_3)\left(\frac{1\text{mole NH}_3}{17\text{g NH}_3}\right)\left(\frac{5\text{ moles }0_2}{4\text{ moles NH}_3}\right)$$

$$\left(\frac{6.02 \times 10^{23}\text{ molecules }0}{1\text{ mole }0_2}\right) =$$

$$3.01 \times 10^{24}\text{ molecules}$$

4. $$\frac{2.0\text{ g}}{182.2\text{g}} = .011\text{ moles Ca}_3\text{P}_2$$

$$.011\left[\text{Ca}_3\text{P}_2 + 6\text{H}_2\text{0}\right] \rightarrow .011\left[3\text{Ca (0H)}_2 + 2\text{PH}_3\right]$$

$$= (.022\text{ moles PH}_3)\left(\frac{33.94\text{g PH}_3}{\text{mole PH}_3}\right)$$

$$= .74\text{g PH}_3$$

LESSON 13

STUDENT ASSIGNMENT SHEET

For Conservation of Mass in Chemical Systems III

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 14. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 13 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 15.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
   - A. Chapter 12, Sections 12.1
   - B. Chapter 3, p. 82 (limiting reagent)
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science, Chapter 12.

PROBLEMS 1. Chapter 12, 12.1, 12.3; 3.11; 3.77

TITLE Problems and the Conservation of Mass II

LESSON NUMBER 14

TIME 50 minutes

DATE PREPARED March 2, 1981

BEHAVIORAL OBJECTIVES As a result of attending this lecture and completing the associated activities, students should be able to:

I Solve quantitative problems based upon balanced chemical equations

I CONDITIONS The conditions under which students must demonstrate minimum competency with the objective are:

A. Using class notes
B. Using class handouts
C. Using the class text
D. By the time Examination Two is administered

II CRITERIA The level reflecting minimum competency with the objective is:

A. Students must be able to answer correctly 90% of all questions related to the objectives of this lesson.

REFERENCES

I INSTRUCTOR
A. Brown & LeMay sample exercises 3.13, 3.15, 3.14
B. Students Guide to Brown & LeMay, Chapter 3
C. Answers to Problems in Brown & LeMay, problems 3.45, 3.69

II STUDENTS
A. Both A and B above

MATERIALS/EQUIPMENT

I INSTRUCTOR
A. Chalk
B. Chalkboard
C. Periodic table
D. One copy of Student Handout 14-1, Using Balanced Equations in Solving Quantitative Problems for Each Student

II STUDENTS
A. Paper
B. Periodic Table

INTRODUCTION Once we are sure an equation is balanced, it may be used to solve any number of chemical problems. Today we shall consider some of those types of problems.

PROCEDURE I LECTURE

A. Problem Types

1. Mole relations in a reaction (mole-mole problems)

   a. Example & steps:

   (1) $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

   (2) relationships from the balanced euqation

      (a) 6 moles CO_2 required to produce 1 mole $C_6H_{12}O_6$

      (b) 6 moles H_2O required to produce 1 mole of $C_6H_{12}O_6$

      (c) 6 moles CO_2 required to produce 6 moles O_2

      (d) 6 moles H_2O required to produce 6 moles O_2

   (3) Problem - If 0.327 moles of $C_6H_{12}O_6$ is formed by a reaction of CO_2 and H_2O, how many moles of CO_2 would be needed?

      (a) Conversion factor

      1 mole $C_6H_{12}O_6$ is equivalent to 6 moles CO_2, so

$$\frac{6 \text{ moles } CO_2}{1 \text{ mole } C_6H_{12}O_6} = 1$$

      (b) $? \text{ moles } CO_2 = .327 \text{ mole } C_6H_{12}O_6 \times \frac{6.00 \text{ moles } CO_2}{1.00 \text{ mole } C_6H_{12}O_6} = 1.96 \text{ mole}$ CO

   (4) Problem - How many moles of O_2 are formed if 0.349 moles of CO_2 is consumed?

      (a) 6 moles CO_2 is equivalent to 6 moles O_2, so $\frac{6 \text{ moles } O_2}{6 \text{ moles } CO_2} = 1$

      (b) $? \text{ moles } O_2 = 0.349 \text{ mole } CO_2 \times \frac{6 \text{ moles } O_2}{6 \text{ moles } CO_2} = 0.349 \text{ moles } O_2$

2. Mass to Mass problems

   a. Example and steps

      (1) $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$

      (2) $264.1g\ CO_2 + 108.1g\ H_2O \rightarrow$

      $180.2g\ C_6H_{12}O_6 + 191g\ O_2$

      (3) 264.1g CO_2 is equivalent to 108.1g H_2O and 264.1g CO_2 is equivalent to 192g O_2 and so on.

      (4) Problem - How many grams of O_2 can be obtained from 65.0g of CO_2?

         (a) $\frac{192g\ O_2}{264.1g\ CO_2} = 1$

$$b.\,gO_2 = 65.0g\ CO_2 \times \frac{192g\ O_2}{264.1g\ CO_2} = 47.2g\ O_2$$

(5) Problem - How many grams H_2O are needed to make 110g $C_6H_{12}O_6$?

$$(a)\ ?g\ H_2O = 110g\ C_6H_{12}O_6 \times \frac{108.1g\ H_2O}{180.2g\ C_6H_{12}O_6} = 66g\ H_2O$$

II ACTIVITY

A. Divide the class in groups of 4

B. Have students solve the probelms listed on Student Handout 14-1.

C. Have students begin probelms 3.45 and 3.69.

APPLICATION I. Students apply skills gained in this lecture to:

A. The activity session described above
B. Homework
C. Future work in chemistry
D. Lesson 15

SUMMARY At the end of the session, summarize briefly the steps involved in analyzing and solving problems related to balanced chemical equations.

EVALUATION Students are evaluated to determine whether they can demonstrate minimum competency with the objectives by evaluation of the Examination Two results.

HOMEWORK I Finish the work started above during the activity session.

II Complete problems 3.45 and 3.69

III Do the required reading for lesson 15 (listed on the course sullabus).

LESSON #14

STUDENT HANDOUT 14-1

Using Balanced Equations in Solving Quantitative Problems

DIRECTIONS:

Solve the following problems using the same method as was used in the lecture. For each problem, complete steps 1-4 below before attempting to actually solve the problem. The reaction used here is that used in the Haber Process, a method od preparing ammonia.

1. The equation $N_2 + 3H_2 \rightarrow 2NH_3$

2. Steps 1-1

   a. What is the first thing to do (make sure the equation is balanced).

   b. What is asked?

   c. With what do you begin?

   d. Derive the conversion factor that is needed.

3. Questions

   a. How many moles of H_2 (dihydrogen) are required to produce 0.889 moles of NH_3 (ammonia).

   b. How many moles of ammonia could be obtained in the Haber process by completely reacting 0.369 moles of dinitrogen?

   c. How many moles of dinitrogen are required to react completely with 0.577 moles of dihydrogen?

   d. How many grams of ammonia can be obtained by complete reaction of 0.267 mole of dinitrogen?

   e. How many grams of H_2 are required to produce 0.275 mole of ammonia.

   f. How many moles of dihydrogen are required to produce a kilogram of ammonia?

   g. How many grams of ammonia could be produced by a complete reaction of 66.6g of H_2?

   h. How many grams of H_2 are required for a complete reaction of 88.8g of N_2?

   i. Which if any is the limiting reagent in each of the problems above? On what was your decision based?

LESSON 14

STUDENT ASSIGNMENT SHEET

For Problems and the Conservation of Mass I

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 15. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 15 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 16.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
   - A. Chapter 3, Sample exercises 3.5; 3.6; 3.7; 3.8; 3.9; 3.10; 3.11
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science, Chapter 3.

PROBLEMS 1. Chapter 3; 3.21; 3.28; 3.30

TITLE Balancing Chemical Equations I

LESSON NUMBER 15

TIME 50 minutes

DATE PREPARED March 4, 1981

BEHAVIORAL OBJECTIVES As a result of attending this lecture and completing the associated activities, students should be able to:

I Relate electron gain or loss to oxidation of reduction

II Define ocidation and reduction

III Balance simple chemical equations by inspection

I CONDITIONS The conditions under which students must demonstrate minimum competency with the objectives are:

A. Using class notes
B. Using class handouts
C. Using the class text
D. By the time Examination Two is administered

II CRITERIA The level reflecting minimum competency with the objectives is:

A. Students must be able to answer correctly 90% of all questions related to the objectives of this lesson.

REFERENCES

I INSTRUCTOR

A. Brown & LeMay
    1. Chapter 3, Section 3.2
    2. Sample exercise 3.1
    3. p. 554
    4. Chapter 7, Section 7.10

B. Students Guide to Brown & LeMay, Chapter 3.
C. Answers to Problems in Brown & LeMay, problems 3.2; 3.3; 3.4; 3.7; 3.56; 7.36; 7.37

II STUDENTS A. Both A and B above

MATERIALS/EQUIPMENT

I INSTRUCTOR

A. Chalk
B. Chalkboard
C. Periodic table
D. One copy of Student Handout 15-1, Balancing Chemical Equations by inspection for each student

II STUDENTS

A. Paper
B. Periodic table

INTRODUCTION

Today we begin the first of three lectures dealing with balancing chemical equations. Once balanced, the equation is an extremely valuable tool. When an equation is encountered, the first thing one should do is determine whether or not the equation is balanced.

PROCEDURE

LECTURE

A. Electron gain and loss in chemical reactions.

1. Compound Formation

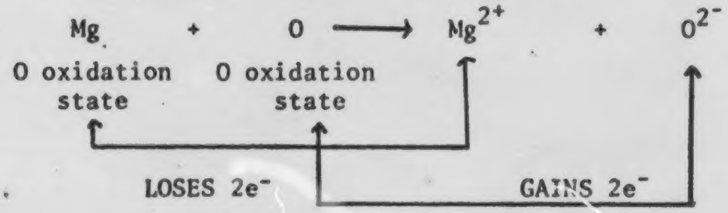
(a) $Mg \longrightarrow Mg^{2+} + 2e^-$
(loss of 2 electrons)

$Mg^{2+} + 2e^- + OH^- \quad Mg^{2+}(OH)^{2-}_2$

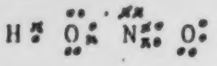
(1) Mg gives up $2e^-$ and valence shell becomes $2s^22p^6$

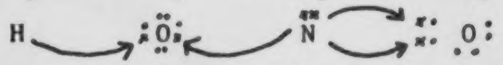
(2) each OH^- picks up one of the e^-

(b) Example



(1) Oxidation numbers on the periodic table indicate the number of e^- gained or lost

(2)  HNO_2

FAMILY	1a	6a	5a	6a
OXIDATION #	1^+	2^-	3^+	2^-
				
THEORETICAL VALENCE POPULATION	$1s^2$	$2s^22p^6$	$2s^22p^6$	$2s^22p^6$
PERIOD	1	2	2	2

(3) Additional examples of e^- transfer:

$$\overset{0}{Na} + \overset{0}{Cl} \longrightarrow \overset{1^+}{Na}\overset{1^-}{Cl}$$

$$\overset{0}{Ca} + \overset{0}{O} \longrightarrow \overset{2^+}{Ca}\ \overset{2^-}{O}$$

$$\overset{0}{Ba} + \overset{0}{F_2} \longrightarrow \overset{2^+}{Ba}\ \overset{-1(2)}{F_2}$$

(4) For rules concerning oxidation numbers, see Student Handout 15-1.

C. Examples of Oxidation Numbers

(1) H_2SiO_3 - What is the charge on Si?

$H_2\ Si\ O_3$

$2(^+1)+3(^-2)$ = The inverse of the charge on Si

$+2 + ^-6$ = Inverse

-4 = Inverse

Charge on Si = $^+4$

becasue $+4 + +2 + -6 = 0$

B. Oxidation and Reduction

1. Definitions

a. Oxidation - a chemical change in which electrons are lost by an atom or group of atoms (an increase in the oxidation number).

$Na \longrightarrow Na^+ + 1e^-$

$0 \longrightarrow {^+1}$

Na is oxidized

b. Reduction - a chemical change in which electrons are gained by an atom or group of atoms (a decrease in oxidation number)

$Cl + 1e^- \longrightarrow Cl^-$

$0 \longrightarrow {^-1}$

Cl is reduced

c. Example of oxidation and reduction

$$\overset{0}{S} + \overset{^+1\,^+5\;3(^-2)-6}{HNO_3} \longrightarrow \overset{^+4\;2(^-2)-}{SO_2} + \overset{4+2-2}{NO} + H_2O$$

S is oxidized from 0 to $^+4$

N is reduced from $^+5$ to $^+2$

C. Balancing by inspection (a trial and error method)

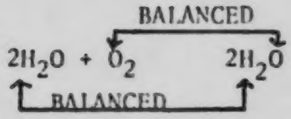
1. Examples

a. $H_2 + O_2 \longrightarrow 2H_2O$

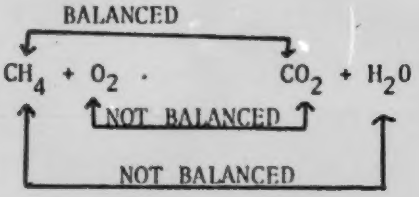
BALANCED

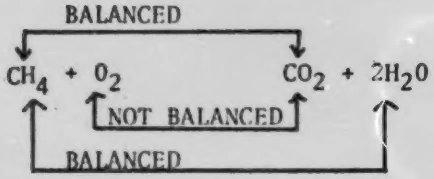
$H_2 + O_2 \longrightarrow 2H_2O$

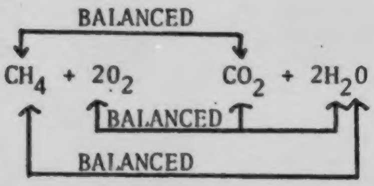
NOT BALANCED


BALANCED
$2H_2O + O_2$
$2H_2O$
BALANCED

b. $CH_4 + O_2$ $CO_2 + H_2O$


BALANCED
$CH_4 + O_2$
$CO_2 + H_2O$
NOT BALANCED
NOT BALANCED


BALANCED
$CH_4 + O_2$
$CO_2 + 2H_2O$
NOT BALANCED
BALANCED


BALANCED
$CH_4 + 2O_2$
$CO_2 + 2H_2O$
BALANCED
BALANCED

II ACTIVITY

A. Divide the class in groups of four
B. Have each group work on the problems presented on Student Handout 15-1.
C. If time permits, have students begin work on homework problems 3.2; 3.3; 3.4; 3.7; 3.56; 7.36; 7.37

APPLICATION

I Students apply understanding gained here to:

A. Lesson 18
B. Chemistry laboratory work
C. Examination 2

SUMMARY

At the end of the lesson, review briefly the definitions presented above.

EVALUATION

Students are evaluated to determine whether they can demonstrate minimum competency with the objectives of the lesson by reviewing the results of Examination 2.

HOMEWORK

I. Complete the exercise started above during the activity portion of the lesson

II. Complete problems 3.2; 3.3; 3.4; 3.7; 3.56; 7.36; 7.37

LESSON #15

Student Handout 15-1

Balancing Chemical Equations by Inspection.

RULES: The following rules apply to the assignment of oxidation numbers.

1. The oxidation state of a free and uncombined element is zero.
2. The oxidation state of hydrogen in compounds is usually +1, except in the case of the metallic hydrides where it is -1.
3. The oxidation state of oxygen in compounds is usually -2, except in peroxides where it is -1 or in fluoride compounds where it may be positive.
4. The algebraic sum of the positi-e and negative oxidation states of all atoms in a neutral molecule is equal to zero.
5. The algebraix sum of the positive and negative oxidation states of all atoms in an ion is equal to the charge of the ion.

DIRECTIONS: Answer each of the following questions as best you can.

1. Given that the oxidation states of hydrogen, oxygen and fluorine respectively are +1, -2 and -1, determine the oxidation states of each of the elements in the following compounds.

   PH_3, H_2S, CrF_3, H_2SO_4, H_2SO_3, $A\ell_2O_3$, $K_4P_2O_2$, $NaAuC\ell_4$, $Ca(C\ell O_2)_2$

2. Balance the following equation by inspection, using the scheme presented during the lecture portion of lesson 15.

   a. $C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$

   b. $H_2SO_4 + 4NaOH \rightarrow H_2O + Na_2SO_4$.

   c. $CuS + HNO_3 \rightarrow Cu(NO_3)_2 + S + H_2O + NO$

   d. $Fe + H_2CO_3 \rightarrow Fe(HCO_3)_2 + H_2$

LESSON 15

STUDENT ASSIGNMENT SHEET

FOR PROBLEMS AND THE CONSERVATION OF MASS II

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 16. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 15 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 17.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
    - A. Chapter 3, Simple exercises 3.13; 3.14; 3.15
    - B. Chapter 20, Introduction, pp. 592-3
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science, Chapter 3.

PROBLEMS 1. Chapter 3; 3.45; 3.69

TITLE Examination 2 Review

LESSON NUMBER 16

TIME 50 minutes

DATE PREPARED March 9, 1981

BEHAVIORAL OBJECTIVES As a result of completing this activity, students should be able to:

I Answer correctly 90% of all questions on Examination Two.

I CONDITIONS The conditions under which students must demonstrate minimum competency with the objective are:

A. Using class notes
B. Using class handouts
C. Using the class text
D. Using the laboratory manual
E. By the time Exam 2 is administered

II CRITERIA Minimum competency is demonstrated when students answer correctly 90% of all questions on Examination Two.

REFERENCES

I INSTRUCTOR A. All of the Brown & Lemay references between Exam 1 and this review

B. All Hill references between Exam 1 and this review.

C. All student handouts between Exam 1 and this review.

II STUDENTS A. A, B and C above
B. All problems solved between Exam 1 and this review.

MATERIALS/ EQUIPMENT

I INSTRUCTOR A. Chalk
B. Chalkboard
C. One copy of Student Handout 16-1, Examination Two Review for each student.

II STUDENTS A. Paper

INTRODUCTION This is a review session. The questions included in Student Handout 16-1 were designed to help students meet and demonstrate minimum competency with course concepts presented between Examination 1 and this review.

PROCEDURE I. ACTIVITY

A. Divide students into groups of four

B. Have students work cooperatively to solve problems listed in Examination Two Review.

APPLICATION Students apply the background gained in this review to Examination Two.

SUMMARY At the end of this session, review briefly the relationship between the questions posed in Student Handout 16-1 and the objectives of lessons presented previously.

EVALUATION Student competency with the objective of this lesson is evaluated by examining the results of Examination Two.

HOMEWORK Continue to prepare for Examination Two.

LESSON # 16
Student Handout 16-1

Examination 2 Review

DIRECTIONS: The following were designed around the course behavioral objectives encountered between examination 1 and this review. They are intended to help you review for examination 2. You should set up and solve each problem exactly as if it were a real exam.

1. If the molecular formula for a sugar is $C_6H_{12}O_6$, what is its empirical formula?

2. The following compound is found in the buildup of boiler sludge,

   $Ca_5(OH)(PO_4)_3$; (a) How many moles of each element would you expect to find in a mole of the sludge ingredient; (b) what is the mass of one mole of the sludge ingredient and; (c) how many atoms of each element are there in 6.02×10^{23} molecules of the sludge ingredient?

3. In the following chemical equation, assume we are dealing with moles. How many moles of each molecule are present?

   $$3\,MgC\ell_2 + 2Na_3PO_4 \rightarrow Mg_3(PO_4)_2 + 6\,NaC\ell$$

4. In question 3, how could you convert from moles of molecules present to molecules present?

5. If, in the boiler related reaction in question 3, 3 moles of $MgC\ell_2$ were in the reaction vessel and mass was to be conserved in the reaction; (a) how many grams of Na_3PO_4 must be add; how many grams of $Mg_3(PO_4)_2$ would be produced and; (c) how many grams of $NaC\ell$ would be produced?

6. The following reaction is part of a metal loss calculation scheme used when dealing with boilers.

   $$3\,Fe(OH)_2 \rightarrow Fe_3O_4 + 2H_2O + H_2$$

   Based upon this calculation; (a) what is the molarily of $Fe(OH)_2$ in the balanced equation assuming the volume of solution to be 1ℓ; (b) What is the normality of a 1ℓ solution of $Fe(OH)_2$; (c) What is the mass of one equivalent weight of $Fe(OH)_2$; (d) how many grams of Fe_3O_4 are produced from 3 moles of $Fe(OH)_2$?

7. Hydrazine, N_2H_4, a powerful reducing agent used in controlling steam plant degradation enter the following reaction with O_2

   $$N_2H_4 + O_2 = 2\,H_2O + N_2$$

   Based upon this oxygen scavenging equation; (a) how many grams of O_2 are removed from boiler water by 20g of N_2H_4?

8. In metal loss calculations, we know that

   $$3\,Fe + 4\,H_2O \rightarrow Fe_3O_4 + 4H_2$$

If 32kg of H_2 gas are produced in a boiler each day, iron is lost in that boiler?

9. The following reactions are some of those which take place during the process of corrosion control in marine steam systems. Balance them by inspection

   a. $Fe_3O_4 + O_2 \rightarrow Fe_2O_3$

   b. $Fe_2O_3 + N_2H_4 \rightarrow Fe_2O_4 + H_2O + N_2$

   c. $CuO + N_2H_4 \rightarrow Cu_2O + H_2O + N_2$

   d. $N_2H_4 + \Delta \rightarrow NH_3 + N_2 + H_2$

   e. $H_2O + CO_2 + MgO \rightarrow Mg(HCO_3)_2$

LESSON 16

Student Assignment Sheet

For Balancing Chemical Equations I

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 17. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 16 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 18.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
   - A. Chapter 3, Section 3.2
   - B. Chapter 7, Section 7.10
   - C. Sample Exercise 3.1
   - D. p. 554
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science, Chapter 3

PROBLEMS

1. Chapter 3; 3.2; 3.3; 3.4; 3.7; 3.56
2. Chapter 7; 7.36; 7.37

TITLE	Examination 2
LESSON NUMBER	17
TIME	50 minutes
BEHAVIORAL OBJECTIVE	As a result of taking this examination students should be able to: I. Identify conceptual weaknesses they have with course material covered between Exam 1 and Exam 2.
I CONDITIONS	Following the correction of the exam.
II CRITERIA	None
REFERENCES	
I INSTRUCTOR	None
II STUDENT	A. Class notes B. Class handouts C. Class text D. Laboratory Manual E. Other material as desired by individual students
MATERIALS/ EQUIPMENT	
I INSTRUCTOR	A. One copy of Exam 2 for each student B. One periodic table for each student
II STUDENTS	A. Both A and B above
INTRODUCTION	Refer students to the instructions on page 1 of the exam.
PROCEDURE	Administer the exam.
APPLICATION	Students use insights gained from the examination to correct misconceptions they may have regarding material covered between Exams 1 and 2 of the course
SUMMARY	None
EVALUATION	None
HOMEWORK	Students do the required reading for lesson 18 as outlined on the course syllabus.

Ch 1 GENERAL CHEMISTRY

Examination 2

Name:____Date:__________________________

DIRECTIONS: Answer each question in the space provided. If there is more than one step involved in the solution, show each step and number the steps successively. It is a good idea to answer first all of the questions you are sure about and then ponder the rest. You have 50 minutes to complete the exam. You may use text books, class notes, laboratory manuals and so on as you work the various problems.

1. A lead nitrate solution, $Pb(NO_3)_2$ is reacted with concentrated hydrochloric acid. The reaction produces a white water-insoluble preciptate. Give correct molecular formula for the precipitate.

Answer__________________

2. What is the law of definite proportions (define it).

3. If 3 g of $Cu(NO_3)_2$ is obtained from allowing 2.0g of Cu to react with excess HNO_3, what is the percent yield of the reaction.

4. A trisodium compound used in boiler feed water treatment has the following molecular formula:

$$(CH_2)_2N(COONa)_3$$

What is the mass of one gram molecular weight of the compound?

5. Water is often called hard when it contains $Mg(HCO_3)_2$. If

$$H_2O + 2CO_2 + MgO \rightarrow Mg(HCO_3)_2,$$

and mass is conserved, how many moles of $Mg(HCO_3)_2$ are produced if 2 moles of CO_2 are present?

Answer__________________

6. In a typical reaction between an acid and a base

$$2\ NH_4OH + H_2SO_4 \rightarrow 2H_2O + (NH_4)_2SO_4$$

If the equation is balanced

a. What molarity acid solution ($H_2 \cdot SO_4$) must be reacted with the base?

Answer ____________________

b. What is the mass of one equivalent weight of the base?

Answer ____________________

c. What is the normality of the acid?

Answer ____________________

d. What is the molarity of the acid?

Answer ____________________

7. Sometimes sodium hydroxide is added to boiler feedwater to remove magnesium chloride. In the reaction,

$$2\ NaOH + MgC\ell_2 \rightarrow 2\ NaC\ell + Mg(OH)_2$$

an insoluble precipitate is formed. What is the mass of $Mg(OH)_2$ formed when 80g NaOH (assume masses of the elements were rounded off) is reacted with one mole of $MgC\ell_2$?

Answer ____________________

8. Balance the following corrosion related equation.

$$Na_2FeO_2 + 4H_2O \rightarrow 6NaOH + Fe_3O_4 + H_2$$

Answer ____________________

9. Define the avogadro number.

10. In a ships boiler, under the influence of heat

$$3\ N_2H_4 \rightarrow 4NH_3 + N_2$$

If the coefficients of reactant and products are in moles, how many atoms of each element are present. Answer this question by filling the table below.

REACTANT		PRODUCTS		
N_2H_2		NH_3		N_2
N	H	N	H	N
Atoms				

LESSON 17

Student Assignment Sheet

For Examination 2 Review

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 18. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 17 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 19.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
   - A. All reading assignments between Exams 1 and 2

PROBLEMS 1. Review textbook problems solved between Exams 1 and 2

TITLE Balancing Chemical Equations II

LESSON NUMBER 18

TIME 50 minutes

BEHAVIORAL OBJECTIVES As a result of attending this lecture and completing the associated activities, students should be able to:

I Identify oxidizing and reducing agents in a chemical equation

II Describe the relationship between electronegativity and oxidation-reduction

III Balance chemical equations by the oxidation number method

I CONDITIONS The conditions under which students must demonstrate minimum competency with the lesson objectives are:

A. Using class notes
B. Using class handouts
C. Using the class text
D. By examination 3

II CRITERIA Minimum competency with the objectives is demonstrated when students are able to:

A. Answer correctly 90% of all lesson related questions

REFERENCES

I INSTRUCTOR A. Brown & LeMay

1. p. 201, review section
2. p. 197, sample exercise 7.11
3. p. 197, sample exercise 7.12

B. Kolb, D. The Chemical Equation Part II: oxidation-reduction reactions. Journal of Chemical Education, 326-331, 55(5), 1978.

C. Myers, R.T. Electronegativity, bond energy, and chemical reactivity. Journal of Chemical Education, 711-712, 56(11), 1979.

D. Vanderwerf, C.A. A consistent treatment of oxidation-reduction. Journal of Chemical Education, 547-551, 25(October), 1948.

E. Yalman, R.G. Writing oxidation-reduction equations. Journal of Chemical Education, 215-218, 36(5), 1959

F. Lockwood, K.L. Redox revisited. Journal of Chemical Education, 326-329, 38(6), 1961.

G. Morris, K.B. the balancing of Oxidation-Reduction Equations. Journal of Chemical Education, 538-540, 15 (November), 1938.

H. Kolb, D. More on balancing redox equations. Journal of Chemical Education, 181-184, 56(3), 1979.

I. Goodstein, M.P. Interpretation of Oxidation-Reduction. Journal of Chemical Education, 452-457, 47(6), 1970.

J. Vanderwerf, C.A., Davidson, A.W. and Sisler, H.H. Oxidation-reduction: A Reevaluation. Journal of Chemical Education, 450-456, 22(September),1946.

K. Answers to Problems in Brown and LeMay for problems 7.38 and 7.39.

II STUDENTS A. A, B, and I above

MATERIALS/ EQUIPMENT

I INSTRUCTOR

A. Chalk
B. Chalkboard
C. One copy of Student Handout 18-1, Balancing Chemical Equations by the Oxidation Number Method, for each student

II STUDENTS A. Paper

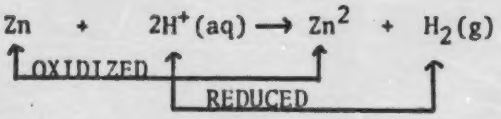
INTRODUCTION Equations cannot always be balanced by observation or trial and error because the process is too complex. One method which attempts to counter the complexity is balancing by the oxidation number method.

PROCEDURE LECTURE

A. Oxidizing and reducing agents in a chemical equation.

1. Definitions

   a. Oxidizing agent - that element in a chemical reaction which is reduced
   b. Reducing agent - that element in a chemical reaction which is oxidized
   c. Example:



$$Zn + 2H^{+}(aq) \longrightarrow Zn^{2} + H_2(g)$$

      (1) Zn is the reducing agent; it reduces H^{+} as it is oxidized
      (2) H^{+} is the oxidizing agent; it oxidizes An as it is reduced

B. The relationship between electronegativity and oxidation-reduction

1. Use p. 193 of student text--table of electronegativities

   a. The lower the electronegativity the better an element is as a reducing agent.

      (1) gives up e^{-} with relative ease

   b. The greater the electronegativity the better an oxidizing agent an element is.

   c. The greater the charge on an element (example Cu^{+1} and Cu^{+2}) the higher its electronegativity and the greater its ability to oxidize.

C. Oxidation number method of balancing equations

1. Step 1 - Identify the compounds that are oxidized and reduced first.

2. Step 2 - Balance elements that are neight oxidized or reduced.

3. Step 3 - Balance oxygen if present

4. Step 4 - Check the balance of hydrogen.

5. Example:

a.

$$\overset{+1}{K}\ \overset{+7}{Mn}\ \overset{\overset{4(-2)}{-8}}{O_4} + \overset{+1}{H}\ \overset{-1}{Cl} \qquad \overset{+2}{Mn}\ \overset{\overset{2(-1)}{-2}}{Cl_2} + \overset{0}{Cl_2} + \overset{+1}{K}\ \overset{-1}{Cl} + \overset{\overset{2(1)}{+2}}{H_2}\ \overset{-2}{O}$$

(1) Step 1 - Assign oxidation numbers

(a) assume H is $^{+}1$ and O is $^{-}2$
(b) elements in the elemental state have a 0 oxidation number

(2) Step 2 - Look for the electron gain and loss

$$\overset{+}{K}\ \underbrace{\overset{+7}{Mn}\ \overset{-8}{O_4} \qquad\qquad \overset{+2}{Mn}}_{\text{Mn gains } 5e^-}\ \overset{-2}{Cl_2}$$

$$\overset{+1}{H}\ \underbrace{\overset{-1}{Cl} \qquad\qquad\qquad \overset{0}{Cl_2}}_{\substack{\text{Cl loses } 1e^- x2 \text{ Cl atoms} \\ 2e^- \text{ loss}}}$$

(3) Step 3 - Electron gain must equal electron loss

i Mn must show $5e^- x2 = 10$ electron gain

ii Cl must show $2e^- x5 = 10$ electron loss

(4) Step 4 - Multiplication factors become molecular coefficients

$$2KMnO_4 + HCl \longrightarrow 2MnCl_2 + 5Cl_2 + KCl + H_2O$$

(5) Step 5 - Balance K

2K in reactants
2K in products -2KCl

(6) Step 6 - Balance Cl

16 Cl in products
16 Cl in reactants 16 HCl

(7) Step 7 - Balance 0

8 0 in reactants
8 0 in products 8 H_2O

(8) Check balance of H

16 H in reactants
16 H in products

b. Balanced equation

$$HKMnO_4 + 16HCl \longrightarrow 2MnCl_2 + 5Cl_2 + 2KCl + 8H_2O$$

6. Additional examples on Student Handout 18-1.

II ACTIVITY

A. Divide students into groups of four
B. Students balance equations and answer questions on Student Handout 18-1.
C. Students complete homework probelms 7.38, 7.39

APPLICATION I. Students apply understanding gained in this lab to:

A. Lesson 19
B. Laboratory activities in Ch 1
C. Future work in chemistry

SUMMARY At the end of the class session, review briefly the steps involved in balancing equations using the oxidation number method.

EVALUATION Students are evaluated to determine whether they can demonstrate minimum competency with the objectives by reviewing the results of Exam 3.

HOMEWORK
I. Complete the balancing zctivity started above
II. Complete homework problem 7.38 and 7.39
III. Accomplish the required reading for Lesson 19 (on course syllabus)

Lesson 18
Student Handout 18-1

BALANCING CHEMICAL EQUATIONS
BY THE OXIDATION NUMBER METHOD

DIRECTIONS: Several examples of balancing quations by the oxidation number method are provided below. Following these examples you will find several equations. You should; (1) identify the oxidizing and reducing agents in each of these equations and; (2) balance each equation using the oxidation number method.

EXAMPLES:

1. 0 +1 +5 -6 +4 -4 +2 -2 +2 -2
$S + HNO_3 \rightarrow SO_2 + NO + H_2O$

S loses 4 e-
N gains 3 e-

A. Multiplication Factors

4e- X 3 = 12e-

3e- X 4 = 12e

B. Coefficients of oxidizing and reducing agents

$3S$, $3SO_2$, $4HNO_3$, $4NO$

$3S + 4HNO_3 \quad 3SO_2 + 4NO + H_2O$

C. Balance Oxygen

Reactants - 12 oxygens

Products - 10 oxygents $2\,H_2O$ = 12

D. $3S + 4HNO_3 \rightarrow 3SO_2 + 4NO + 2H_2O$

E. Check balance of hydrogen

Reactants - 4Hydrogens

Products - 4 Hydrogens

2. +1 +5 -6 0

$HNO_3 + I_2 \rightarrow HIO_3 + NO_2 + H_2O$

N gains 1e-

I loses 5e- X 2 I atoms = 10e-

A. Multiplication factors

N -5le - $^{x}10$ = 10e- so

10 HNO_3 and $10NO_2$

I -5e X 2 atoms = 10e- so

2 HIO_3

B. Balance Oxygen

$10HNO_3 + I_2 \quad 2HIO_3 + 10NO_2 + H_2O$

Reactants -30 oxygens
Products - 2x3=6 oxygens in
HIO_3; 10x2=20 oxygens in NO_2;
total 20+6=26 present; therefore $4H_2O$

$10 HNO_3 + I_2 \rightarrow 2HIO_3 + 10NO_2 + 4H_2O$

C. Check balance of hydrogen

Reactants -10 hydrogens

Products - 2 hydrogens +8 hydrogens

3. $C_3H_8O + CrO_3 + H_2SO_4 \rightarrow$

$Cr_2(SO_4)_3 + C_3H_6O$

A. Steps

(1) Write the correct products for all reactants and products

(2) Assign oxidation to the atoms in the equation

(3) Identify atoms which change oxidation number

(4) Make the number of atoms that change oxidation number the same on both sides of the equation by inserting temporary coefficients

(5) Complete the total change in oxidation number for the oxidation-reduction to occur

(6) Make the electrron loss equal the electron gain by multiplication using appropriate factors

(7) Balance the remainder by inspection.

B. Solution

(1) $C_3H_8O + CrO_3 + H_2SO_4 \quad Cr_2(SO_4)_3 + C_3H_6O + H_2O$

(2) -2 +1 -2 +6 -2 +1 -2 +3 -2 -2/3 +1 -2

$C_3H_8O + CrO_3 + H_2SO \rightarrow Cr_3(SO_4)_3 + C_3H_6O + H_2O$

a. SO_4^{2-} appears on both sides of the equation, therefore, the net oxidation number is used

(3) Atoms that change

a. carbon goes from -2 to -4/3

b. chromium goes from +6 to +3

(4) -2 +6 +3 -4/3

$1C_3H_8O + 2CrO_3 + H_2SO_4 \rightarrow 1Cr_2(SO_4)_3 + 1C_3H_6O + H_2O$

a. both formulas containing carbon have three carbon atoms; the temporary coefficients of C_3H_8O and C_3H_6O are 1.

b. the $Cr_2(SO_4)_3$ contains two Cr while CrO_3 contains one. The temporary coefficient of CrO_3 therefore is 2.

(5) Change in oxidation numbers (total)
2e- loss
6e- gain
3x(2)=-6 2x(+6)=+12 2x(+3)=+6 3x($^{-4}$/3)=-4

$$\overset{-2}{1C_3H_8O} + \overset{+6}{2CrO_3} + H_2SO_4 \rightarrow \overset{+3}{1Cr_2(SO_4)_3} + \overset{-4/3}{1CrH_{36}O} + H_2O$$

(6) Electron gain and loss

a. carbon (2e-loss) x3=6e-loss

b. chromium (6e-loss) x1=6e-loss

c. factors 3 and 1 are used to multiply the temporary coefficients

$3C_3H_8O + 2CrO_3 + H_2SO_4 \rightarrow 1Cr_2(SO_4)_3 + 3C_3H_6O + H_2O$

3x1 1x2 1x1 3x1

(7) Inspection

a. 3 SO_4 in the product -3 SO_4 needed in reactants, therefore 3 H_2SO_4

b. 30 hydrogens in the reactants; 18 hydrogens in acetone thus 12 more are needed, therefore $6H_2O$.

(8) Balanced Equation

$$3C_3H_8O + 2\,CrO_3 + 3H_2SO_4 \rightarrow Cr_2(SO_4)_3 + 3C_3H_6O + 6H_2O$$

PROBLEMS: Balance the following equations using the oxidation number method.

1. $KC\ell + MnO_2 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + C\ell_2 + H_2O$
2. $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NH_4NO_3 + H_2O$
3. $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$
4. $SO_2 + HNO_3\ H_2O \rightarrow H_2SO_4 + NO$
5. $I_2 + NaOH \rightarrow NaI + NaIO_3 + H_2O$

TITLE Balancing Chemical Equations III

LESSON NUMBER 19

TIME 50 minutes

DATE PREPARED March 16, 1981

BEHAVIORAL OBJECTIVES As a result of attending this lecture and completing the associated activities, students should be able to:

I Balance chemical equations by the ion electron method;

II Balance chemical equations by the algebraic method

I CONDITIONS The conditions under which students must demonstrate minimum competency with the lesson objectives are:

A Using class notes
B Using class handouts
C Using the class text
D By Examination 3

II CRITERIA Minimum competency with the objectives of the lesson is demonstrated when a student:

A Answers correctly 90% of all lesson related questions

REFERENCES

I INSTRUCTOR

A. Brown & LeMay
  1. Chapter 19, Section 19.1
  2. p. 557, Sample exercise 19.1

B. Kolb, D., The Chemical Equation, Part II: Oxidation Reduction Reactions. *Journal of Chemical Education*, 326-331, *55(5)*, 1978.

C. Vanderwerf, C.A., Davidson, A.W., and Sisler, H.H., Oxidation-Reduction: A Re-Evaluation. *Journal of Chemical Education*, 450-456, *22(September)*, 1945.

D. Steinbach, O.F., Non-Stoichiometric Equations, *Journal of Chemical Education*, 66-69, *21(January)*, 1944.

E. Hill, Chapter 19

F. Answers to Problems in Brown & LeMay, problems 19.1, 19.3, 19.4

MATERIALS/EQUIPMENT

I INSTRUCTOR

A. Chalk
B. Chalkboard
C. One copy for each student of Student Handout:
  1. 19-1, Balancing Chemical Euqation by the Ion-Electron Method
D. Periodic table

II STUDENTS
A. Paper
B. Periodic table

INTRODUCTION

Today we will discuss two additional methods by which chemical equations may be balanced; the ion-electron method and the algebraic method

PROCEDURE

I. LECTURE

A. The ion-electron method (sometimes called the half-reaction method).

1. Deals with oxidation-reduction equations;
2. Involved writing individual half-reactions;
3. Requires starting with the ionic form of the equation;
    a. look for the electrolytes--they ionize in solution
4. $S + HNO_3 \longrightarrow SO_2 + NO + H_2O$
    $S + H^+ + NO_3^- \longrightarrow SO_2 + NO + H_2O$
    a. half reactions

        Oxidation: $S \longrightarrow SO_2$
        Reduction: $NO_3^- \longrightarrow NO$

    b. Balance half reaction
        (1) oxidation--oxygens not balanced
            $2H_2O + S \longrightarrow SO_2$
        (a) hydrogens not balanced
            $2H_2O + S \longrightarrow SO_2 + 4H^+$
        (b) balance charges by adding electrons to one side or the other
            $2H_2O + S \longrightarrow SO_2 + 4H^+ + 4e^-$
        (2) reduction--handled in same manner as oxidation
            $NO_3^- \longrightarrow NO + 2H_2O$
            $4H^+ + NO_3^- \longrightarrow NO + 2H_2O$
            $3e^- + 4H^+ + NO_3^- \longrightarrow NO + 2H_2O$
    c. Complete half reactions
        $2H_2O + S \longrightarrow SO_2 + 4H^+ + 4e^-$
        $3e^- + 4H^+ + NO_3^- \longrightarrow NO + 2H_2O$
    d. Electron loss in the oxidation step must equal electron gain in the reduction step
        (1) the electron terms must cancel each other--there must be 4 X 3 = 12e⁻ on each side

            Oxidation: $3\,[2H_2O+S] \longrightarrow 3\,[SO_2+4H^++4e^-]$

            $6H_2O+3S \longrightarrow 3SO_2+12H^++12e^-$

            Reduction: $4\,[3e^-+4H^++NO_3^-] \longrightarrow 4\,[NO+2H_2O]$

            $12e^-+16H^++4NO_3^- \longrightarrow 4NO+8H_2O$
    e. Cancellation Step
        (1) add the two half reactions
            $6H_2O+3S \longrightarrow 3SO_2+12H^++12e^-$
            $\underline{+12e^-+16H^++4NO_3^- \longrightarrow 4NO+8H_2O}$
            $12e^-+6H_2O+16H^++3S+4NO_3^- \longrightarrow 3SO_2+12H^++4NO+12e^-+8H_2O$
        (2) follow rules of mathematics to cancel terms, keeping all signs positive

(a) $^{-}12e^{-}+12e^{-}+6H_2O+16H^{+}+3S+4NO_3^{-} \Rightarrow$

$3SO_2+12H^{+}+4NO+8H_2O+12e^{-}-12e^{-}$

(b) $-6H_2O+6H_2O+16H^{+}+3S+4NO_3^{-} \Rightarrow$

$3SO_2+12H^{+}+4NO+8H_2O-6H_2O$

(c) $-12H^{+}+16H^{+}+3S+4NO_3^{-} \Rightarrow$

$3SO_2+4NO+2H_2O+12H^{+}-12H^{+}$

(d) $3S+4NO_3^{-}+4H^{+} \quad 3SO_2+4NO+2H_2O$

B. The Algebraic Method of balancing redox equations

1. $S + HNO \longrightarrow SO_2 + NO + H_2O$

a. Assign arbitrary leters to represent the unknown coefficients of the reactants and products.

$xS + YHNO_3 \longrightarrow ZSO_2 + pNO + qH_2O$

(1) sulfur $x=z$
(2) nitrogen $y=p$
(3) hydrogen $y=2q$
(4) oxygen $3y=2z+p+q$

By substitution $3y=2z+y+y/2$

$$2z=3y-y-y/2$$
$$2z=3y/2$$
$$4z=3y$$
$$y=3/4\ z$$

clearing the fraction, if $y=4$, $z=3$ and it follows that $x=3$, $p=4$ and $q=2$

b. Substitute the derived values in the original equation.

$xS + yHNO_3 \longrightarrow zSO_2 + pNO + qH_2O$

$3S + 4HNO_3 \longrightarrow 3SO_2 + 4NO + 2H_2O$

II. ACTIVITY

A. Divide students in groups of four
B. Students balance equations listed on Student Handout 19-1
C. Students solve homework problems, 19.1, 19.3 and 19.4

APPLICATION

I. Background gained in this lesson is applied to:
A. Chemistry lab work
B. Future work in chemistry

SUMMARY

At the end of the session, review briefly the steps in solving redox equations by the ion-electron method.

EVALUATION

Students are evaluated to see if they can demonstrate minimum competence with the objectives by reviewing the results of Examination 3.

HOMEWORK

I Complete the work started in the activity above
II Complete homework problems 19.1, 19.3, 19.4
III Do reading for lesson 20 (on course syllabus)

LESSON #19
STUDENT HANDOUT 19-1

Balancing Chemical Equations by the ion-electron (half reaction) method

RULES: The following rules apply to balancing redo equations using the ion-electron method.

1. Write the equation in ionic form and separate it into two half reactions.

2. Balance each half reaction separately, adding H 0 where needed to balance oxygen atoms and H+ where needed to balance hydrogen atoms.

    A. For alkaline solutions use OH^- to balance oxygen and H 0 to balance hydrogen.

3. Add electrons to one side of each half-reaction so as to balance the charges.

4. Make electron charges equal by multiplying one or both half-reactions by appropriate numbers.

5. Add the two half-reactions together, elminating all duplication.

EXAMPLES:

1. $HNO_3 + H_2S \rightarrow NO + S + H_2O$

$H^+ + NO_3^- + H_2S \rightarrow NO\ S + H_2O$

$NO_3^- \rightarrow NO$

$H_2S \rightarrow S$

$NO_3^- \rightarrow NO + 2H_2O$

$4H+ + NO_3^- \rightarrow NO + 2H_2O$

$3e^- + 4H^+ + NO_3^- \rightarrow NO + 2H_2O$

$H_2S \rightarrow S + 2H^+$

$H_2S \rightarrow S + 2H^+ + 2e^-$

$2\left[3e^- + 4H^+ + NO_3^-\right] \rightarrow 2\left[NO + 2H_2O\right]$

$6e^- + 8H^+ + 2NO_3^- \rightarrow 2NO + 4H_2O$

$3\left[H_2S\right] \rightarrow 3\left[S + 2H^+ + 2e^-\right]$

$3H_2S \rightarrow 3S + 6H^+ + 6e^-$

$6e^- + 8H^+ + 2NO_3^- + 3H_2S \rightarrow 2NO + 4H_2O + 3S + 6H^+ + 6e^-$

$-6e^- + 6e^- + 8H^+ + 2NO_3^- + 3H_2S \rightarrow 2NO + 4H_2O + 3S + 6H^+ + 6e^- - 6e$

$-6H^+ + 8H^+ + 2NO_3^- + 3H_2S \rightarrow 2NO + 4H_2O + 3S + 6H^+ - 6H^+$

$2H^+ + 2NO_3^- + 3H_2S \quad 2NO + 3S + 4H_2O$

2. $Zn + NaNO_3 + NaOH \rightarrow Na_2ZnO_2 + NH_3 + H_2O \quad NO_3^- \rightarrow NH_3$

   a. In alkaline solutions, each excess oxygen atom is balanced by adding one OH^- to the same side and one H_2O to the opposite side.

   b. Each excess hydrogen atom is blanced by adding one OH^- and one H_2O to the opposite side

$NO_3^- + 3H_2O \rightarrow NH_3 + 6OH^-$

$NO_3^- + 3H_2O + 3H_2O \rightarrow NH_2 + 6OH^- + 3OH^-$

(1) $3OH^-$ on the right and $3H_2O$ on the left balance the excess hydrogen of NH_3.

$NO_3^- + 6H_2O \rightarrow NH_3 + 9OH^-$

$NO_3^- + 6H_2O + 8e^- \rightarrow NH_3 + 9OH^-$

$Zn \quad ZnO_2^{2-}$

$4OH^- + Zn \rightarrow ZnO_2^{2-} + 2H_2O$

$4OH^- + Zn \rightarrow ZnO_2^{2-} + 2H_2O + Ze$

$1\left[NO_3^- + 6H_2O + 8e^-\right] \rightarrow 1\left[NH_3 + 9OH^-\right]$

$4\left[4OH^- + Zn\right] \rightarrow 4\left[ZnO_2^{2-} + 2H_2O + 2e^-\right]$

$16OH^- + 4Zn \rightarrow 4Zn_2^{2-} + 8H_2O + 8e^-$

$NO_3^- + 6H_2O + 8e^- + 16OH^- + 4Zn \rightarrow$

$NH_3 + 9OH^- + 4Zn_2^{2-} + 8H_2O + 8e^-$

$NO_3^- + 4Zn + 7OH^- \rightarrow NH_3 + 4ZnO_2^{2-} + 2H_2O$

$NaNO_3 + 4An + 7NaOH \rightarrow NH_3 \; 4\,Na_2ZnO_2 + 2H_2O$

PROBLEMS: Balance the following problems using the ion-electron method.

1. $K^+MnO_4^- + H^+Cl^- \rightarrow K^+Cl^- + Mn^{2-}(Cl^-)_2 + H_2O + Cl_2$
2. $CuS + H^+NO_3^- \rightarrow Cu^{2-}(NO_3^-)_2 + S + H_2O + NO$

Lesson #19

3. $H^+NO_3^- + H^+I^- \rightarrow NO + I_2 + H_2O$

4. SO_4^{2-} (aq)→ SO (g) (acidic solution)

LESSON 19

Student Assignment Sheet

For Balancing Chemical Equations II

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 20. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 19 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 21.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
   - A. p. 201, review section
   - B. Chapter 7, Sample exercises 7.11; 7.12
2. Kolb, D. the chemical equation part II: oxidation-reduction reactions. Journal of Chemical Education, 326-331, 55(5). 1978

PROBLEMS 1. Chapter 7, 7.38, 7.39

TITLE Chemical States of Matter

LESSON NUMBER 20

TIME 50 minutes

DATE PREPARED March 16, 1981

BEHAVIORAL OBJECTIVES As a result of attending this lecture and completing the associated activities, students should be able to:

I Describe time temperature curves
II Compute heats of fusion and vaporization based upon empirical data

I CONDITIONS The conditions under which students must demonstrate minimum competence with the objectives are:

A. Using class notes
B. Using class handouts
C. Using the class text
D. By the time Examination 3 is administered

II CRITERIA The level reflecting a demonstration of minimum competency with the objective is:

A. Students must answer correctly 90% of all questions related to the objectives of this lesson

REFERENCES

I INSTRUCTOR

A. Brown & LeMay
  1. Chapter 9, Section 9.1
  2. Chapter 4, Section 4.3
  3. Chapter 18, Sections 18.2, 18.3
  4. p. 538, Figure 18.5
  5. p. 335, Figure 11.29
  6. Chapter 11, Sections 11.1; 11.2; 11.3; 11.8

B. Solutions to Exercises in Chemistry the Central Science, problems 11.10; 11.12; 11.18; 11.20; 11.56; 11.58; 11.77

C. Hill, Chapter 11

II STUDENTS A. Both A and C above

MATERIALS/ EQUIPMENT

I INSTRUCTOR

A. Chalk
B. Chalkboard
C. One copy for each student of Student Handouts:

1. 20-1, Problems Dealing With the States of Matter;
2. 20-2, Phases of Matter;
3. 20-3, Comparison of Time-Temperature Curves for Pure Substances and Mixtures

II STUDENTS

A. Paper

INTRODUCTION

When dealing with chemical systems there are three states with which we must be conversant; solids, liquids and gases. The state of a substance at any given moment depends upon several variables. We shall talk about variables for the remainder of this lecture.

PROCEDURE

I LECTURE

- A. Time-temperature curves and states of matter (pure substances)
    - 1. Use student handout 20-2, Phases of Matter for discussion purposes
        - a. Key to 20-2:
            - (1) AB--solid phase
                - (a) definite shape
                - (b) slightly compressible
                - (c) molecular vibration occurs within a definite area
                - (d) energy input causes temperature rise
            - (2) T_f--the melting or freezing temperature
            - (3) BC--solid/liquid phase
                - (a) H_{fusion}
                - (b) energy input goes to causing a change of state but not temperature increase
            - (4) CD--liquid phase
                - (a) decrease in density over solid phase
                - (b) slightly compressible
                - (c) no specific shape
                - (d) specific volume
                - (e) energy input causes temperature rise
                - (f) viscosity (internal resistance to flow) depends upon nature of the substance
                - (g) molecular activity is random
            - (5) T_b--boiling temperature
            - (6) DE--liquid-gas phase
                - (a) $H_{vaporization}$
                - (b) energy input goes to causing a change of state (breaking intermolecular bonds) not to a temperature change
            - (7) EF--gas phase
                - (a) no specific shape
                - (b) no specific volume--expands to fill the space available
                - (c) decrease in volume over the liquid
                - (d) highly compressible
                - (e) energy input causes increase in temperature

B. Time-temperature curves (mixtures and pure substances)
1. Use studnet handout 20-3, Comparison of Time and Temperature Curves for Pure Substances and Mixtures for discussion purposes
a. Key to 20-3
(1) Mixture--composed of more than one substance
(2) Substances have different melting and boiling points
b. Causes of melting and boiling points
(1) atmospheric pressure
(2) intermolecular bonding energies
c. Computation of energy requirements to cause a temperature increase in a substance
a. Units
(1) molar heat (enthalpy change) of fusion (ΔH_{fusion}) measured in KJ/mol
(a) Example: -H_2O
ΔH_{fusion}=6.02KJ/mol or 1435 cal/mol

$6.02KJ+H_2O(S) \longrightarrow H_2O(l)$ *18g)
ENDOTHERMIC RXN

$\Delta H_{crystallization}$=-6.02KJ/mol
EXOTHERMIC RXN $H_2O(l) \longrightarrow H_2O(S)$
-sign means it gives up that much heat/mol
(2) molar heat (enthalpy change) of vaporization
$\Delta H_{vaporization}$=9713 cal/mol (at 100°C)
(1 cal=4.184J)

$H_2O(\ell) \longrightarrow H_2O(g)$ (18g)

(3) heat capacity the amount of energy required to raise a substance a degree C
(a) measured in J/°C-g
$H_2O(S)$=2.03
$H_2O(\ell)$=4.184
(4) normal boiling point--the temperature at which the vapor pressure of the liquid equals one standard atmosphere (760 m m hg).
(5) vapor pressure--the pressure created by a gas over a liquid-caused by molecules with high kinetic energies breaking into the gas phase

II ACTIVITY

A. Divide the class in groups of 4 students each
B. Have students complete the activity on student handout 20-1
C. Begin homework problems 11.10; 11.11; 11.18; 11.20; 11.56; 11.58; 11.77

APPLICATION Students apply skill and background gained here to:

A. Lab work
B. Future work in the maritime industry

SUMMARY At the end of the period summarize briefly the differences between solids, liquids and gases

EVALUATION Students are evaluated to determine whether they can demonstrate minimum competency with the objectives of the lesson by reviewing the results of examination three

HOMEWORK

I Complete the activity started above

II Complete homework problems 11.10; 11.12; 11.18; 11.20; 11.56; 11.58; 11.77

III Do the required reading for lesson 21 (listed on the course syllabus

LESSON #20

STUDENT HANDOUT 20-1

Problems Dealing With the States of Matter

DIRECTIONS: Answer the following questions to the best of your ability.

1. Describe the relationship between Figure 11-3 on page 307 of your text and the states of matter.

2. Describe the relationship between Figure 11.5 on page 310 of your text and the states of matter.

3. Describe the relationship between Figure 11.6 on page 311 of your text and the states of matter.

4. If the heat of vaporization for $C\ell_2$ is 20.4KJ/moℓ how much energy is required in the following reaction (assume coefficients are in moles).

   $20\ C\ell_2(e) \rightarrow 20\ C\ell_2(g)$

5. Assume H_2O is a solid at 0^0C. How much energy is required in the following reaction.

   $15\ H_2O\ (s) \rightarrow 15H_2O\ (g)$

6. The heat capacity of H_2O is 4.184 $J/^0C$-g and 18 moles of $_0H_2O$ at 37^0C are mixed with a substance having a heat capacity of 0.85 $J/^0C$-g.

   If$_0$the temperature of the substance and the temperature of the water fell 20^0C before the two substances reached a uniform temperature, how much of the unknown was present?

7. The specific heat of the corrosion product

   Fe_2O_3 is 0.151 cal/g^0C.

   a. How much heat is required to raise the temperature of 200g of $Fe\ O_3$ from 20.0^0C to 30.0^0F.

   b. What is the molar heat capacity of Fe_2O_3?

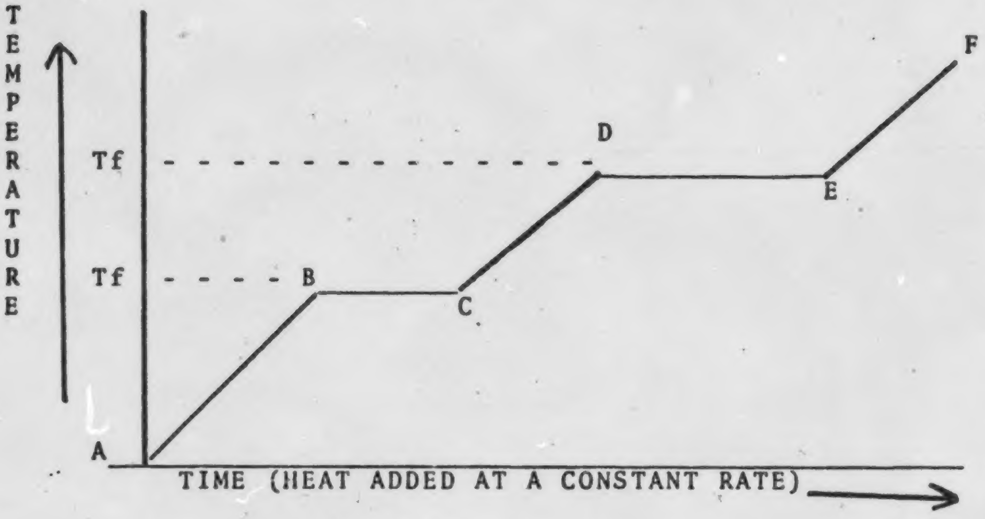
LESSON #20

STUDENT HANDOUT 20-2

Phases of Matter

Time Temperature (Heating) Curves.

Time temperature curves present a graphic picture of what happens to a pure substance at constant atmospheric pressure as heat is added at a constant rate.

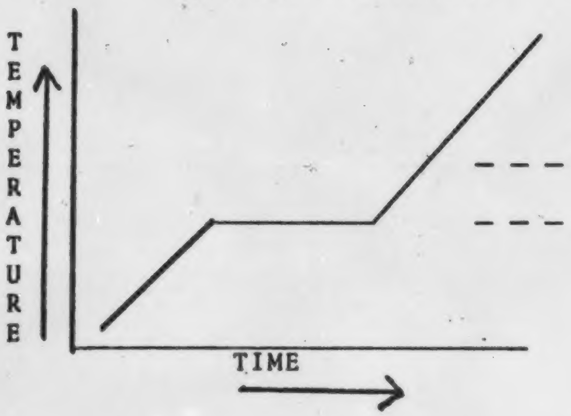


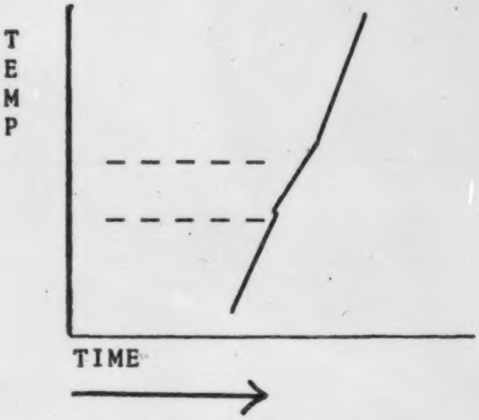
LESSON #20
STUDENT HANDOUT 20-3

Comparison of time temperature curves for pure substances and mixtures.

DISCUSSION: When compared, the solid-liquid and the liquid-gas phases of pure substances should be flat while the same phases for a mixture should be sloped.

PURE SUBSTANCES





LESSON 20

Student Assignment Sheet

For Balancing Chemical Equations III

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 21. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 20 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 22.

READING ASSIGNMENT

1. Brown and LeMay *Chemistry: The Central Science*
   - A. Chapter 19, Section 19.1, Sample exercise 19.1
2. Kolb, D. the chemical equation part II: oxidation-reduction reactions. *Journal of Chemical Education*, 326-331, *55(5)*, 1978
3. Hill *Student Guide to Brown and LeMay, Chemistry: The Central Science*, Chapter 19

PROBLEMS 1. Chapter 19; 19.1; 19.3; 19.4

TITLE The Gaseous State and Kinetic Molecular Theory

LESSON NUMBER 21

TIME 50 minutes

DATE PREPARED March 17, 1981

BEHAVIORAL OBJECTIVES As a result of participation in this lecture and completion of the associated activites, students should be able to:

I. Describe the Avogadro hypothesis

II Solve Dalton's, Boyle's Charles' and Ideal Gas Law problems

I CONDITIONS The conditions under which students must demonstrate minimum competency with the objectives of the lesson are:

A. Using class handouts
B. Using class notes
C. Using the class text
D. By the time examination 3 is administered
E. After having solved the problems listed below for this lesson

II The level reflecting a demonstration of minimum competency with the objectives of this lesson is:

A. Students must answer correctly 90% of all objective related questions when exposed to them

REFERENCES

I INSTRUCTOR
A. Brown & LeMay
  1. Chapter 9, sections 9.2, 9.3, 9.4, 9.5
B. Answers to Problems in Chemistry: the Central Science, for problems 9.3, 9.8, 9.13, 9.14, 9.18
C. Hill, Chapter 9

II STUDENTS A. Both A and C above

MATERIAL/ EQUIPMENT

I INSTRUCTOR
A. Chalk
B. Chalkboard
C. Periodic table
D. One copy of Student Handout 21-1, Gas Law Problems, for each student

II STUDENT
A. Paper
B. Periodic Table

INTRODUCTION Gases as we have mentioned before have a list of characteristics all their own. These characteristics and their associated laws apply regardless of whether the gas is oxygen, carbon dioxide or steam. They have no definite shape or volume and in fact expand to fill the space available. Their mixtures are homeogeneous and their pressures the sum of pressures of the individual constituents.

Several laws and at least one hypothesis are important in understanding the physical and chemical action of gases. They are the Avogadro hypothesis, Dalton's Law, Boyle's Law, Charles' Law and the Ideal Gas Law. These laws and hypotheses are discussed in the following context.

PROCEDURE I LECTURE

A. Units of measurements used with gases
   1. P = pressure
   2. T = temperature
   3. S = standard
   4. STP = standard temperature and pressure
      a. P = 1 atmosphere=760 m m ng (1 m m ng = 1 TORR)
      b. T = $0^{o}C = 273.15^{o}K^{-}$ use $^{o}K^{-}$ division by zero is undefined
   5. N = The number of atoms, molecules or moles of molecules
   6. R = Gas law constant
      a. Numerical values on p. 250 of Brown & LeMay

B. Ideal Gas Law PV = nRT
   1. $P = \frac{nRT}{V}$

$$P = \frac{n\ (\cancel{\text{moles}})\ \frac{R(\cancel{\ell}\text{-atm})}{\cancel{\text{moles}}\text{-}\cancel{K}}\ T(^{o}\cancel{K})}{V(\cancel{\ell})}$$

$$P = \frac{nRT}{V}\ \text{atmosphere}$$

   a. $R = 0.08205\ \frac{\ell\text{-atm}}{\text{mole-K}}$

C. Avogadro Hypothesis
   1. Based upon the Ideal Gas Law - one mole of any gas at ST Poccypies 22.4ℓ
      a. At STP $V_1=V_2=V_3$ //ℓ

         (assume 1, 2, 3 all are 1 mole)

         $V_N = V_{H2} = V_H$ etc.

D. Dalton's Law for nonreacting gases (gases are homogeneous)
   1. $P_{total} = P_1 + P_2 + P_3$
   2. The total pressure exerted by a gas is the simple sum of the partial pressures of the individual constituents.

E. Temperature-Volume Law-(Charle' Law)-the volume of a given sample of gas kept at constant pressure is directly proportional to its Kelvin temperature
   1. $V \propto T$ (fixed mass and pressure)
   2. $V = CT$ $\quad \frac{V}{T} = C$
      a. C = constant of proportionality
   3. Using 1 and 2 as times, since

      $\frac{V_1}{T_1} = C$ and $\frac{V_2}{T_2} = C$

      $\frac{V_{time_1}}{T_{time_1}} = \frac{V_{time_2}}{T_{time_2}}$ (as long as mass and pressure do not change)

      $V_2 = V_1 \left(\frac{T_2}{T_1}\right)$

      $T_2 = T_1 \left(\frac{V_2}{V_1}\right)$

F. Pressure-Volume Law (Boyles Law)-At a constant temperature the volume of a given sample of gas is inversely proportional to the applied pressure

   (T and mass held constant)

   1. $V \propto \frac{1}{P}$

   2. $V = \left(\frac{1}{P}\right) C$

   3. $V_1P_1=C$ and $V_2P_2=C$

   4. $V_1P_1 = V_2P_2$

   5. $P_2=P_1 \left(\frac{V_1}{V_2}\right)$

      $V_2=V_1 \left(\frac{P_1}{P_2}\right)$

II ACTIVITY

A. Divide the class in groups of four
B. Have each group set up and solve problems listed on Student Handout 21-1
C. Begin homework problems 9.3, 9.8, 9.13, 9.14, 9.18

APPLICATION I. Students apply skills gained here to:
A. Lesson 22
B. Laboratory Activities
C. Course work in thermodyanmics
D. Future work in chemistry

SUMMARY	At the end of the session, review briefly the laws and hypotheses presented above
EVALUATION	Students are evaluated to determine whether they can demonstrate minimum competence with the lesson objectives by reviewing the results of Exam 3.
HOMEWORK	I Complete the activity started above II Complete homework problems 9.3, 9.8, 9.13, 9.14, 9.18 III Do the required reading for lesson 22 (listed on the course syllabus

LESSON 21
STUDENT HANDOUT 21-1

Gas Law Problems

DIRECTIONS: Apply the laws presented in this lecture and answer the following questions.

1. If 10 moles of H_2O are converted to steam at 400°F for use in a boiler tube whose volume is 3 liters, how much pressure (in torr) must the boiler tube walls be able to withstand?

2. Analysis of steam in a ships evaporator shows steam pressure to be 4 atm, dissolved oxygen pressure to be .8 torr and dissolved CO_2 pressure to be 1.3 atm. What pressure must the evaporator walls be able to withstand?

3. In a winch engine,

   $$2\ C_8H_{18}(g) + 25\ O_2(g) \rightarrow 16CO_2\ (g) + 18H_2O\ (g)$$

   How many liters of O_2 must be added at STP for every 2 moles of octane burned?

4. What is the new pressure on the boiler tube wall in question #1 if the steam temperature increases to 300°C?

5. If the volume of a one kiloliter acetelene tank normally under 5 atm of pressure is suddenly decreased to 3/4 of its original volume by a collision, what pressure must the walls withstand if explosion is to be averted?

LESSON 21

Student Assignment Sheet

For Chemical States

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 22. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 21 Problem Set (lsited on separate pages) must be completed and submitted at the beginning of Lecture 23.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
    - A. Chapter 9; Section 9.1
    - B. Chapter 4, Section 4.3
    - C. Chapter 18, Sections 18.2; 18.3, Figure 18.5
    - D. Chapter 11, Sections 11.1; 11.2; 11.3; 11.8, Figure 11.29

PROBLEMS 1. Chapter 11, 11.10; 11.12; 11.18; 11.20; 11.56; 11.58; 11.77

TITLE Interrelationships of the Gas Laws

LESSON NUMBER 22

TIME 50 minutes

DATE PREPARED March 17, 1981

BEHAVIORAL OBJECTIVES As a result of participation in this lecture and completion of the associated activities, students should be able to:

I. Solve gas law problems

I CONDITIONS The conditions under which minimum competency with the objectives of this lesson must be demonstrated are:

A. Using class notes
B. Using class handouts
C. Using the class text
D. By examination three
E. Regardless of whether it is volume, temperature or pressure that is manipulated

II CRITERIA The level of understanding which reflects a demonstration of minimum competency with the lesson objective is:

A. Students, when asked, must answer correctly 90% of all objective related questions

REFERENCES

I INSTRUCTOR

A. Brown & LeMay
   1. All sections covered for lecture 20
B. Hill, Chapter 9
C. Solutions to Problems in Chemistry: The Central Science, Problems 9.70; 9.21; 9.23; 9.27; 9.29; 9.15

II STUDENTS A. Both A & B above

MATERIALS/ EQUIPMENT

INSTRUCTOR

A. Chalk
B. Chalkboard
C. Periodic table
D. One copy of Student Handout 22-1, Gas Law Problems II, for each student

STUDENT

A. Paper
B. Periodic table

INTRODUCTION

This lecture is an extension of Lesson 21. Today we look critically at the interrelationships of the gas laws. You must be able to solve gas law problems regardless of the parameter which varies. This lesson should increase your skills in that regard.

PROCEDURE

I. LECTURE

A. Ideal Gas Law

1. $PV=nRT$ & $\frac{PV}{T} = nR$

   a. With n constant (no change in no. of moles)
   b. R = a constant
   c. 1 & 2 are times 1 & 2, places 1 & 2, situations 1 & 2 and so on

2. $$\frac{P_1V_1}{T_1} = Rn = \text{constant}$$

   $$\frac{P_2V_2}{T_2} = Rn = \text{constant}$$

3. $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

4. or if n_1 and n_2 are different from time-to-time, etc.

   $$\frac{P_1V_1}{T_1n_1} = R \quad) \quad \frac{P_2V_2}{T_2n_2} = R$$

   THEREFORE:

   $$\frac{P_2V_2}{T_2n_2} = \frac{P_1V_1}{T_1n_1}$$

B. Boyle's Law - With temperature and number of moles held constant

1. From 4 above in Ideal Gas Laws $\frac{P_1V_1}{\cancel{T_1n_1}} = \frac{P_2V_2}{\cancel{T_2n_2}}$

   because $T_1n_1 = T_2n_2$

   $P_1V_1 = P_2V_2$

C. Charles Law - With pressure and number of moles held constant

1. From 4 in the Ideal Gas Law equation

$$\frac{\cancel{P_1}V_1}{T_1\cancel{n_1}} = \frac{\cancel{P_2}V_2}{T_2\cancel{n_2}}$$

because $n_1 = n_2$ and $P_1 = P_2$

therefore $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

2. From 4 in the Ideal Gas Law with volume and number of moles held constant

$$\frac{P_1\cancel{V_1}}{T_1\cancel{n_1}} = \frac{P_2\cancel{V_2}}{T_2\cancel{n_2}}$$

because $n_1 = n_2$ and $V_1 = V_2$

therefore $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ and $P_1 = \frac{T_1P_2}{T_2}$

D. Gas density from pressure-temperature data

1. $PV = nRt$

2. $\frac{P}{RT} = \frac{n}{V}$ $\left(n = \frac{\text{grams}}{\text{molecular weight}}\right)$

   and $\frac{n}{V} = \frac{\text{mass}}{\text{volume}} = \text{density}$

3. mass (MW) = moles x molecular weight

4. $\frac{MW}{1} \times \frac{n}{V} = \frac{MW}{1} \times \frac{P}{RT} = \text{density}$

   and $\frac{n(\cancel{\text{moles}})}{V(\text{liters})} \times \frac{MW(\text{grams})}{\cancel{\text{moles}}} = \frac{n(MW)}{V}\ g/\ell$

5. $\frac{MW(\text{grams})}{\cancel{\text{mole}}} \times \frac{P(\cancel{\text{atm}})}{R\left(\frac{\text{liter-}\cancel{\text{atm}}}{\cancel{{}^{\circ}K}\text{-}\cancel{\text{mole}}}\right)T^{\circ}K} = \frac{(MW)P}{RT}\ g/\ell$

II. ACTIVITY

A. Divide the class in groups of 4 students each
B. Have each group solve the problems listed on Student Handout 22-1
C. Have students begin homework problems 9.21; 9.23; 9.27; 9.29; 9.15

APPLICATION

I. Students apply understanding gained here to:

A. Laboratory work in chemistry
B. Examination 3
C. Future employment in the maritime industry

SUMMARY At the end of the period, review briefly the interrelationships of the gas laws.

EVALUATION Students are evaluated to determine whether they are able to demonstrate minimum competency with the objectives by reviewing the results of Examination 3.

HOMEWORK

I Complete the activity started above

II Complete homework problems 9.21; 9.23; 9.27; 9.29; 9.15

III Do the required reading for Lesson 23 (listed on the course syllabus

LESSON 22

STUDENT HANDOUT 22-1A

Solutions to Gas Law Problems II

1. a. $^\circ K = (110-32)\ 5/9 + 273 = 316.3^\circ K$

b. $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ but V is constant

c. $\frac{P_1T_2}{T_1} = P_2$; $\frac{760\ \text{mmhg} \times 316.3^\circ K}{273^\circ K}$

$= 880.54$ mmhg $= 1.15$ atm.

d. Nothing happens

2. a. 1140 mmgh = 1.5 atm

b. $\frac{P_1}{T_1} = \frac{P_2}{T_2}$, $T = \frac{P_2T_1}{P_1}$

c. $\frac{1140\ \text{mmhg} \times 273^\circ K}{760\ \text{mmhg}} = 136^\circ C =$

$277.7^\circ F$

3a. MW of $C\ell_2 = 70.9$g/mole

$$d = \frac{\frac{70.9\text{g}}{\text{mole}}\left(\frac{95\ \text{TORR}}{760\ \text{TORR/atm}}\right)}{(.0825\ell\text{-atm/mole-}^\circ K)(273^\circ K)}$$

4a. $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ therefore

$\frac{P_2}{P_1} = \frac{T_2}{T_1}$ and $T_2 = \frac{T_1P_2}{P_1}$

b. $\frac{(2.24\ \text{atm})(760\ \text{torr/atm})(25.0 + 273^\circ K)}{844\ \text{torr}}$

c. $328^\circ C$

5a. $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

b. $V_2 = \frac{P_1V_1T_2}{P_2T_1}$

LESSON #22
STUDENT HANDOUT 22-1

Gas Law Problems II

DIRECTIONS: Using the gas laws and related concepts, solve the following problems.

1. Refrigerated gas is stored at STP in a compartment on a ship. The containers will withstand 1.5 atm. The refrigeration system fails and the temperatures soar to 110°F. What happens?

2. With regard to the containers discussed in question 1. At what temperature F would you expect them to burst?

3. Chlorine gas, $C\ell_2$, was found in the battery storage area of a diesel submarine. What was the density of $C\ell_2$ in g/l if the partial pressure of $C\ell_2$ in the compartment was found to be 95 torr at 0°C.

4. A boiler tube can withstand a pressure of 2.24 atm and 0.2 g of steam exerts a pressure 844 torr at 25°C. What is the maximum temperatures in °C that the tube may be heated without exceeding 2.24 atm?

5. 10ℓ of steam at 400°F exerts a pressure of 25 atm at one point in a ships boiler. At another point the temperature is found to be 600°F and the pressure 20 atm. What volume of steam is found at the second location?

LESSON 22

Student Assignment Sheet

For the Gaseous State and Kinetic Molecular Theory

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 23. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 22 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 24.

READING ASSIGNMENT

1. Brown and LeMay *Chemistry: The Central Science*
    A. Chapter 9; Sections 9.2; 9.3; 9.4; 9.5
2. Hill *Student Guide to Brown and LeMay, Chemistry: The Central Science*, Chapter 9

PROBLEMS

1. Chapter 9; 9.3; 9.8; 9.13; 9.14; 9.18

TITLE Solutions

LESSON NUMBER 23

TIME 50 minutes

DATE PREPARED March 18, 1981

BEHAVIORAL OBJECTIVES As a result of participation in this session and completion of the associated activities, students should be able to:

- I Define terms associated with solutions
- II Describe how solutions of specific strengths may be mixed
- III Describe mathematically how to dilute solutions of known strength
- IV Relate solution strength to chemical reactions

I CONDITIONS The conditions under which students must demonstrate minimum competency with the lesson objectives are:

- A. Using class notes
- B. Using class text
- C. Using class handouts
- D. By the time Examination 3 is administered

II CRITERIA The level reflecting minimum competence with the objectives of this lesson is:

- A. Students must, when asked, answer correctly 90% of all lesson related objectives

REFERENCES

I INSTRUCTOR

- A. Brown & LeMay
    - 1. Chapter 3, Section 3.9
    - 2. Chapter 12, Sections 12.1; 12.2; 12.3; 12.4; 12.7
- B. Hill, Chapter 12
- C. Solutions to Problems in Chemistry: The Central Science for problems 12.1; 12.3; 12.6

II STUDENTS

- A. Both A & B above

MATERIALS/ EQUIPMENT

I INSTRUCTOR

- A. Chalk
- B. Chalkboard
- C. One copy of Student Handout 23-1, Solution Related Problems for each student
- D. Periodic table

II STUDENTS

A. Paper
B. Periodic table

INTRODUCTION

Solutions as we have mentioned previously are composed of a solute and a solvent. The solute is the ingredient in lesser supply while the solvent is the one in greater supply. There exist solid-solid, solid-liquid, liquid-liquid, liquid-gas solutions and so on. Solvents contain solutes in solutes in solution to the point of saturation at which time solute and solvent and solute separate. With liquid-gas solutions, the gas may bubble off. With solids dissolved in solvent, the solids reform in the orignal or a new substance forms as a precipitate and with liquid-liquid solutions, an interface may form between the two liquids.

Solutions are of interest to marine engineers as they may have an effect upon corrosion in marine systems.

PROCEDURE

I LECTURE

A. Solution Strength

1. Molarity = $\frac{\text{moles of solute}}{\text{liters of solution}}$

   a. Steps in mixing a $\underline{1}$ molar SOLN

      (1) Weigh one mole of solute (mass in grams)
         (a) 39.97g NaOH for example
      (2) Add a sufficient quantity of distilled H_2O to make the total volume 1ℓ
      (3) Stir

2. Normality - the number of equivalent weights per liter of solution

   a. Steps in mixing a 1 normal SOLN
      (1) Determine the number of moles of OH^-, H^- or [illegible] contributed by 1 mole of solute.
         (a) $H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$, for example

            so 98g (1 mole) H_2SO_4 gives up 2 moles of H^+
      (2) Determine the mass of one equivalent weight of solution
         (a) In the case of H_2SO_4 98g/2
      (3) To the mass of one equivalent weight, add sufficient distilled H_2O to make a liter of SOLN

B. Dilution Formula

1. $\text{Molarity}_{\text{initial}} \times \text{Volume}_{\text{initial}} = M_{\text{final}} \times V_{\text{final}}$

$$M_I = \frac{M_f V_f}{V_I} \quad , V_I = \frac{M_f V_f}{M_I} \text{ etc.}$$

2. Used when converting solutions of one strength to another

C. Solute Solubility

1. Solution - a chemical system in dynamic equilibrium

a. SOLUTE & SOLVENT $\underset{(B)}{\overset{(A)}{\rightleftarrows}}$ SOLN

(1) A direction, solute dissolves

(a) $BaSO_4(S)+H_2O \longrightarrow$

$Ba^{2+}(aq) + SO_4^{2-}$(aq indicates solution hydration or aquation)

(2) B direction, solute crystallizes and a precipate is formed

(a) $Ba^{2+} + SO_4^{2-} \longrightarrow BaSO_4$

b. Equilibrium Concentration

(1) the amount of solute needed to saturate a solution

(a) Saturation-solvent cannot hold additional solute without a precipitate forming

(b) At saturation, $RATE_A=RATE_B$

(c) Excess solute+SOLN $\longrightarrow$ Saturated SOLN + precipitate

2. Effects on Solubility
   - a. molecular structure of the solvent
   - b. molecular structure of the solute
   - c. temperature
   - d. pressure (gases)

3. Expression of Solubility

a. solubility = $\frac{\text{grams solute}}{\text{100g solvent}}$ at a specific temperature

D. Solution Types

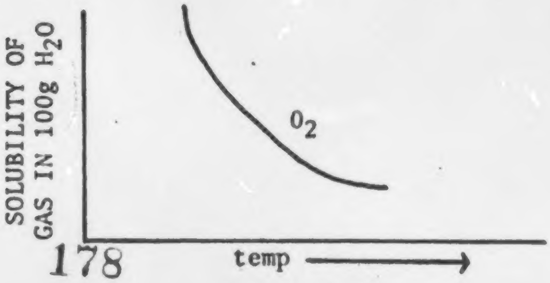
1. Endothermic - solubility of the solute increases as ^{o}K increase (see fig. 12.10, 359 of text)

a. Solute + solvent + heat $\longrightarrow$ SOLN

2. Erothermic - solubility of the solute decreases as temperature increases

a. Solute + solvent $\longrightarrow$ SOLN + heat

b. Example - Gas - Liquid solubility (principle upon which the deaeration heater in a boiler works)



c. $C = K_p$

(1) gas concentration in SOLN=C
(2) K=(Henry's) constant $\left(\frac{\text{mole/liter}}{\text{ATM}}\right)$
(3) pressure=P

II ACTIVITY

A. Divide the class into groups of four
B. Have students begin homework problems 12.1; 12.3; 12.6

APPLICATION I Students apply understanding gained in this session to:
A. Future work in shipboard chemistry
B. Laboratory work in chemistry

SUMMARY At the end of the session, review briefly the basic ingredients of a solution, how solutions are mixed and how they are diluted.

EVALUATION Students are evaluated to determine whether they can demonstrate to minimum competency with the objectives of the lesson by reviewing the results of Examination 3

HOMEWORK I Complete the activity started above
II Complete homework problems 12.1; 12.3; 12.6
III Do the required reading for Lesson 24 (listed on the course syllabus)

LESSON 23

STUDENT HANDOUT 23-1

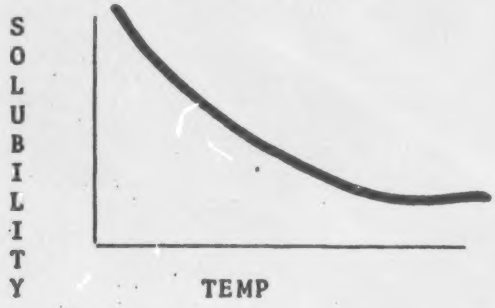
DIRECTIONS: Using the solution related concepts discussed in today's lecture, work with your fellow students to solve the following problems. In some cases, a periodic table is necessary.

1. H_2SO_4 is one of the chemicals used in a chemical company's phenolthalein - alkalinity test of boiler feed water condition. The test requires the use of .IN solution. Your shipboard chemical cabinet, however, contains only 3 liters of .3M H_2SO_4 solution. How can you resolve the issue? Your resolution should be done mathematically as well as verbally.

2. The chloride test is often run on shipboard potable water. The test requires the use of silver nitrate, $AgNO_3$. If we assume that the test solution, $AgNO_3$, must be saturated and the solubility of $AgNO_3$ at 25°C is .155g/mℓ, explain how you would mix a liter of the solution.

3. How would you mix 3ℓ of 0.1N H_2SO_4 for use in a pH of condensate test?

4. One function of a deaeration heater in a marine engineering system is to help limit corrosion by removing dissolved oxygen. If the following is true for O_2:



   a. Is O_2 exothermic or endothermic?

   b. If the solubility of O_2 in 25°C water is 4.89 g/mℓ and at 100°C is 2.46 g/mℓ and the corrosion reaction sequence is:

   $$4\ Fe + 3O_2 \rightarrow 2Fe_2O_3$$

   How many fewer grams of Fe_2O_3 can be formed per liter of hot rather than cold water?

   c. If the boiler is under a pressure of 10 torr, how many moles of O_2 are dissolved per liter solution?

5. In marine engineering plants, scale is formed when the following reaction occurs

Lesson 23 - Student Handout 23-1

$$Ca(HCO_3)_2 \xrightarrow{\Delta} \underline{CaCO_3} + CO_2 + H_2O$$

If a saturated solution at 100°C contains .0019g/mℓ $CaCO_3$

a. How moles of Ca CO_3 are there in 10ℓ of plant water

b. If the solution at 25°C is saturated at .00153g/mℓ, is the solution endothermic or exothermic?

c. If a solution, saturated at 100°C decreases to 25°C how many moles/ℓ of $CaCO_3$ precipitate?

d. In part C, how many grams of Ca $(HCO_3)_2$ could be formed from the precipitate assuring the reaction is reversible?

LESSON 23

Student Assignment Sheet

For Interrelationships of the Gas Laws

DIRECTIONS The reading assignmeht outlined below should be completed prior to attending Lecture 24. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 23 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 25.

READING ASSIGNMENT

1. Brown and LeMay *Chemistry: The Central Science*
    A. Review reading done for the previous lecture
2. Hill *Student Guide to Brown and LeMay, Chemistry: the Central Science*, Chapter 9

PROBLEMS 1. Chapter 9; 9.70; 9.21; 9.23; 9.27; 9.29; 9.15

TITLE	Examination 3 Review
LESSON NUMBER	24
TIME	50 minutes
DATE PREPARED	March 19, 1981
BEHAVIORAL OBJECTIVES	As a result of completing the activity associated with this lesson students should be able to: A. Identify misconceptions they have regarding subjects covered between Examinations 2 and 3.
I CONDITIONS	The conditions under which students must demonstrate minimum competency with the objectives of the lesson are: A. Using class notes B. Using class handouts C. Using the classs text D. By Examination 3
II CRITERIA	Minimum competency with the objectives of this lesson is demonstrated when students: A. Answer correctly 90% of all questions on Examination 3
REFERENCES	
I INSTRUCTOR	None
II STUDENTS	A. All readings in Brown & LeMay assigned between Exams 2 and 3 B. All homeowrk problems assigned between Exams 2 and 3.
MATERIALS/ EQUIPMENT	
I INSTRUCTOR	A. One copy of Student Handout 24-1, Examination 3 Review for each student
II STUDENTS	A. Paper
INTRODUCTION	This is a review session. The questions included in Student Handout 24-1 were designed to help students meet the minimum competency requirement for class sessions conducted between Exams 2 and 3.

PROCEDURE	ACTIVITY A. Divide students in groups of four B. Have students work cooperatively to solve problems included in Student Handout 24-1.
APPLICATION	Students apply experience gained here to the solution of problems included on Examination 3.
SUMMARY	At the end of the class, review briefly the relationship between the questions posed on the Student Handout and the objectives listed for the various lessons between Exams 2 and 3.
EVALUATION	Students' competence with the objective of this lesson is evaluated by examining the results of the results of Exam 3.
HOMEWORK	Continue to prepare for Exam 3.

LESSON 24

STUDENT HANDOUT 24-1

Examination 3 Review

DIRECTIONS: The following group of problems were designed around the behavioral objectives considered during the period between exams 2 and 3. They are intended to help you review the concepts embraced between those exams. You should set up and solve each problem as if it were an actual examination.

1. Elemental copper is in the following corrosion related reaction in the presence of concentrated acid is converted as follows:

$$Cu + HNO_3 \rightarrow Cu\,(NO_3)_2 + NO_2 + H_2O$$

a. Balance the equation using oxidation number method.

b. If 1 Kg of $Cu(NO_3)_2$ is produced in the reaction how much Cu is lost to the engineering parts involved?

2. Unprotected iron engineering parts are converted to Fe $(OH)_2$ even in the absence of dissolved $_2$ or CO_2. The reaction involved is as follows:

$$Fe + H_2O \rightarrow 6Fe(OH)_2 + H_2$$

a. Balance the equation by inspection.

b. How much iron is lost to the engineering plant for each mole of H_2 gas produced?

3. In boiler feedwater treatment, impurities are removed from the water. In one of the treatments, lime, Ca $(OH)_2$ is added to the water and the following reaction occurs:

$$Ca(OH)_2 + Mg\,(HCO_3)_2 \rightarrow Mg\,(OH)_2 + CaCO_3 + H_2O$$

a. Balance the equation by inspection.

b. Determine whether any of the elements are oxidized or reduced.

4. In eletrolylic corrosion, a process which often involves ship hulls, the following occurs:

$$3H_2O + Fe \rightarrow Fe_2O_3 + 6H^+ + 6e-$$

a. Is iron oxydized or reduced?

b. Is iron the oxydizing or reducing agent?

5. Balance the following equation using the ion-electron method

$$CuS + HNO_3 \rightarrow Cu(NO_3)_2 + S + H_2O + NO$$

6. Specific heat is the quantity of heat liberated or given up when the temperature of 1g of a substance rises or falls 1^0C. Based upon the following table;

Substance	Specific heat cal/g-^0C	Freezing Point ^0C	Heat of Fusion cal/g	Boiling Point ^0C	Heat of Vaporization cal/g
H_2O (ℓ)	1.00	0.0	79.7	100.0	540
H_2O (S)	0.46				
NH_3 (ℓ)	1.05	-77.7	79.3	33.4	327
NH_3 (S)	0.50				

a. How much energy is required if 5g of ice at 243 ^0K are to be converted to steam?

b. How much energy is required to convert 5g of NH_3 at -243^0K to a vapor?

c. How much heat must be dissipated by a boiler when 50 g of steam at 373^oK decreases in temperature to 90^oF?

7. When acetylene stored in a tank at 25^oC and 1 atm increases in temperature to 600^0F, what happens to tank pressure and why?

8. Why is the business of shipping LNG so dangerous? Explain in chemical terms.

9. If the volume of a boiler tube suddenly decreases by 40% because of scale build up what happens to steam pressure in the tube if it was originally 600 psi (your results should be mmhg)?

10. What is the effect of increasing the quantity, of oxygen stored in a tank, from two to six moles if the pressure and temperature remain constant?

11. If the temperature of the air in a submarines ballast tank increased by forward starboard 20% but the tank at the same time developed a leak which effectively increased, the volume by 30%, what would the effect be on tank pressure?

12. On the compression cycle in an 4 stroke internal combustion engine, if piston diameter is 4 inches, the stroke length is 4 inches, the pressure on the piston is 1 atm at the bottom of the stroke and the total volume of the cylinder chamber is 10% greater than the void created when the piston at its lowest position, what is the pressure on the piston when it is in the full up position, just before firing occurs.

13. How would you make 4ℓ of 0.2N H_2SO_4 for use in your shipboard boiler water analysis program?

14. What is the final volume of a solution of NaOH if its orginal volume is 2M and the strength must be 0.2N?

15. If the following quation is balanced related to your boiler maintenance

$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$

program and you find 10 mℓ of a sample to contain 0.1M of the acid (H_2SO_4)

How much 1M NaOH must you add per liter of boiler water if mass is to be conserved?

16. If stack gas of a merchant vessel is 600°F and contains the following gases

gas	mass of the gas	volume
CO	60g	10ℓ
CO_2	88g	6ℓ
SO_2	50g	8ℓ

a. What is the total pressure exerted by the gas on the stack walls?

b. What would be the effect on total pressure increasing the number of moles of SO_2 by a factor of 4?

LESSON 24

Student Assignment Sheet

For Solutions

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 25. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 24 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 26.

READING ASSIGNMENT

1. Brown and LeMay *Chemistry: The Central Science*
    - A. Chapter 3; Section 3.9
    - B. Chapter 12; Sections 12.1; 12.1; 12.3; 12.4; 12.7
2. Hill *Student Guide to Brown and LeMay, Chemistry: The Central Science*, Chapter 12

PROBLEMS 1. Chapter 12; 12.1; 12.3; 12.6

TITLE	Examination 3
LESSON NUMBER	25
TIME	50 minutes
DATE PREPARED	March 20, 1981
BEHAVIORAL OBJECTIVES	As a result of taking this examination students should be able to: I. Identify areas of conceptual weakness with course material covered between examinations 2 and 3.
I CONDITIONS	The conditions under which students must demonstrate competency with the objective of this lesson are: A. Using class handouts B. Using class notes C. Using the class text D. Using the class laboratory manual
II CRITERIA	A. None
REFERENCES	
I INSTRUCTOR	A. None
II STUDENTS	A. Brown & LeMay B. Class handouts C. Class notes D. Class laboratory manual
MATERIALS/ EQUIPMENT	
I INSTRUCTOR	A. One copy of Examination 3 for each student B. One periodic table for each student
II STUDENTS	A. Both A and B above
INTRODUCTION	Refer students to the instructions on page 1 of the exam.
PROCEDURE	Administer the exam
APPLICATION	Students use the insight gained from this exam to correct any misunderstandings they may have over concepts covered between exams 2 and 3.
SUMMARY	None
EVALUATION	None
HOMEWORK	Students do the reading required for lesson 25 (outlined on the course syllabus).

LESSON 25

Student Assignment Sheet

For Examination 3 Review

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 26. The problems listed from the class text are intended to support and amplify material presented during the lecture portion of each class session. They should be completed after the class session is over. The problems included in the Lesson 25 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 27.

READING ASSIGNMENT

1. Brown and LeMay *Chemistry: The Central Science*
    A. Review required reading between Exams 2 and 3

PROBLEMS

1. Review textbook problems solved between Exams 2 and 3

EXAM 3 Spring 1981

Ch-1 GENERAL CHEMISTRY

NAME: ______________________________ DATE: ______________

DIRECTIONS: Answer each question in the space provided. If there is more than one step involved in a solution, show each step and number the steps successively. It is a good idea to answer first all questions about which you are sure and then ponder the rest. You may use text books, class notes, laboratory manuals and the like as you work the various problems.

---

1. The chloride ion, Cl^-, is one whose presence in boiler feedwater condensate could lead to engineering plant degradation. Describe one method of evaluating condensate for the presence of the chloride ion.

2. Magnesium iron and zinc are two metals commonly found in ship parts. When reacted with acid we often find the following to occur.

   $Zn + 2HCl \longrightarrow ZnCl_2 + H_2$

   $Fe + 2HCl \longrightarrow FeCl_2 + H_2$

   $Mg + 2HCl \longrightarrow MgCl_2 + H_2$

   If all three metals are in close proximity in an engineering plant and emersed in the same solution (acid water in this case):

   (a) Which would you expect to corrode away first? Ans ____________

   (b) Which would be the second to show the results of corrosion? Ans ____________

   (c) Why did you chose the answers you did for parts (a) and (b)?

3. What would be the freezing point of a solution containing 3.5g of salt (NaCl) dissolved in 100g of water if the normal freezing point of water is 0.0°C? Show all of your work.

4. Above 350°F in boilers there is a direct reaction between iron and water or steam the reaction is:

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

a. Name the oxidizing agent ______________________

b. Name the reducing agent ______________________

c. Name the element being oxidized ______________________

d. Name the element being reduced ______________________

5. General corrosion, characterized by chemical or electro-chemical reaction, occurs uniformly over an exposed surface for a number of reasons. The reaction that occurs is as follows:

$$Fe + O_2 + H_2O \longrightarrow Fe(OH)_2$$

In the space provided, apply the rules for and balance the equation using the oxidation number method. Show all of your work and number each step. Credit will not be given if work is not shown even though the answer may be correct.

6. List the steps involved in balancing chemical equations by the half reaction method.

7. Stack gases from a merchant vessel are analyzed and found to contain the following species: SO_2, CO, and CO_2. The partial pressures of the gases respectively are 3 atm, 1520 mm hg and 2280 torr. What is the total pressure in mm hg exerted on the sides of the stack?

8. Pressurized air is often used for starting shipboard diesel engines. Pressure in the tanks is usually increased by pumping additional quantities of air into such tanks. Suppose the air tank on your ship at 420 K contains 10 moles of air at 760 mm hg but a pressure of 4560 mm hg is required to start an auxiliary diesel generator. How many moles of air must be pumped into the tank before it can serve its intended purpose?

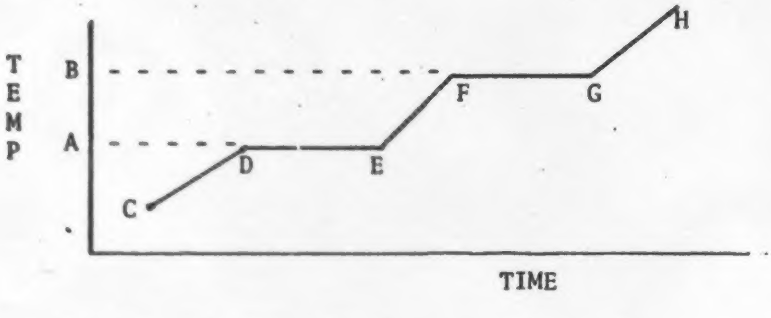
Ans ________________

9. One of the chemicals used in your shipboard water analysis program is 0.1 N H_2SO_4. During rough sea conditions your ship experiences extensive rolling and the only 0.1N H_2SO_4 you have is destroyed. You do however have onboard one liter of one molar H_2SO_4 and adequate supplies of distilled water.

a. How much distilled water must be added to the 1M H_2SO_4 to dilute it to 0.1N H_2SO_4? Ans ______

b. What is the final volume of the H_2SO_4 after the 1M solution has been diluted? Ans ______

10. The following figure is a time-temperature curve for octane.



a. Is C_8H_{18} a pure substance or a mixture? Ans ______

b. Line segment CD probably prepresents? Ans ______

c. A is probably the? Ans ______

d. B is probably the? Ans ______________

e. Line segment DE probably represents? Ans ______________

f. Line segment EF probably represents? Ans ______________

g. Line segment FG probably represents? Ans ______________

h. Line segment GH probably represents? Ans ______________

LESSON #26

STUDENT HANDOUT 26-2

Acid-Base Neutralization Problems

DIRECTIONS: Using the concepts discussed in lesson 26, solve the following problems. Remember there is a difference between the concentration in pH and pOH. When solving the pH problems, refer to Student Handout 26-1, pH and pOH

1. A 100 ml sample is taken from your ships boiler whose capacity 5000ℓ, and analyzed to determine the pH. 10mℓ of 0.1N H_2SO_4 are required to neutralize the sample.

   a. What is the normality of the water in the boiler

   b. What is the pH of the boiler water?

   c. How many moles of OH^- would you expect to find in each liter of the water?

   d. What is the mass of H^+ contained in the boiler water?

2. A phenolthalein alkalinity test is often used to test the quality of marine engineering plant water. How many moles/liter of OH^- are dissolved in the water when the indicator is in the center of its color change range?

3. If,

$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$

. How many liters of the acid (strngth 2M) are required to neutralize 10ℓ of boiler water containing NaOH, if the water strength is 2N?

4. 2 liters of boiler water is found to have an OH^- concentration of 2.0×10^{-6}

   a. What is the pH?

   b. What is concentration of H^+ in moles/liter?

   c. How many grams of H^+ are dissolved in the sample?

5. Based upon sample exercise 15.4 in your class text, is boiler water in which the H^+ concentration is found to be 3.0×10^{-8}M basic or acidic?

6. Using an indicator the center of whose color change range is pOH=7, it takes 15 mℓ of 0.1N H_2SO_4 to neutralize 20 mℓ of boiler water. What is the pH of the boiler water?

7. Water often undergoes a process called autoionization in which the following occurs

$$H_2O + H_2O \rightleftarrows H_3O + +OH^-$$

Because the reaction is reversible each side can be considered products or reactants depending upon the arrow chosen. In each case, identify the acid and the base.

LESSON #26

STUDENT HANDOUT 26-1

pH and oPH

(H+) (MOLE/LITER)	pH	(OH-) MOLE/LITER	pOH	NATURE	
10^{1}	-1	10^{-15}	15	ACID	ACIDITY INCREASE
10^{0}	0	10^{-14}	14		
10^{-1}	1	10^{-13}	13		
10^{-2}	2	10^{-12}	12		
10^{-3}	3	10^{-11}	11		
10^{-4}	4	10^{-10}	10		
10^{-5}	5	10^{-9}	9		
10^{-6}	6	10^{-8}	8		
10^{-7}	7	10^{-7}	7	NEUTRAL	
10^{-8}	8	10^{-6}	6	ALKALINE	ALKALINIT INCREASE
10^{-9}	9	10^{-5}	5		
10^{-10}	10	10^{-4}	4		
10^{-11}	11	10^{-3}	3		
10^{-12}	12	10^{-2}	2		
10^{-13}	13	10^{-1}	1		
10^{-14}	14	10^{0}	0		
10^{-15}	15	10^{1}	-1		

TITLE	Final Examination
LESSON NUMBER	27 and 28
TIME	120 minutes
DATE PREPARED	April 1, 1981
BEHAVIORAL OBJECTIVES	As a result of completing this activity, students should be able to: I Identify course-wide misconceptions they may have regarding basic chemistry
I CONDITIONS	The conditions under which students must demonstrate facility with the objectives of the course are: A. Using class notes B. Using class handouts C. Using the class text D. Using the class laboratory manual
II CRITERIA	None
REFERENCES	
I INSTRUCTOR	None
II STUDENTS	A. Class text B. Class laboratory manual C. Class handouts D. Class notes
MATERIALS/ EQUIPMENT	
I INSTRUCTOR	A. One copy of the final examination for each student present
II STUDENTS	A. Paper B. Periodic table
PROCEDURE	Administer the examination
APPLICATION	Students apply the understanding grained here to future work in chemistry
SUMMARY	None
EVALUATION	None
HOMEWORK	None

LESSON 27

Student Assignment Sheet

For Acids and Bases

DIRECTIONS The reading assignment outlined below should be completed prior to attending Lecture 28. The problems included in the Lesson 27 Problem Set (listed on separate pages) must be completed and submitted at the beginning of Lecture 28.

READING ASSIGNMENT

1. Brown and LeMay Chemistry: The Central Science
   - A. Chapter 15, Sections 15.1; 15.2; 15.3; 15.8
   - B. pp. 67-69
2. Hill Student Guide to Brown and LeMay, Chemistry: The Central Science, Chapter 15
3. Kolb, D. Acids and Bases, Journal of Chemical Education, 459-464, 55(7), 1978
4. Kolb, D. the pH concept. Journal of Chemical Education, 49-53, 56(1), 1979

Ch-1 GENERAL CHEMISTRY
Final Examination

Name ____ Spring 1981

DIRECTIONS: Answer each question in the space provided. If there is more than one step involved in a solution, show each step and number the steps successively. It is a good idea to answer first all questions about which you are sure and then ponder the rest. You may use text books, class notes, laboratory manuals and the like as you work the various problems. You have 2 hours to complete the examination.

1. Define equivalence point and end point:

Equivalence Point ____

____

____

____

____

End Point ____

____

____

____

____

2. The following tetrasodium compound is one used commonly in feedwater treatment.

$$(CH)_2 N_2 (CH_2COONa)_4$$

a. What is the mass of one mole of the compound?

Ans ______________________

b. How many atoms of each element are present in one mole of the compound?

CH_2 N_2 CH_2 $COON_a$

C	H	H	C	H	C	O	Na

c. If water has Kfp of 1.86 ^{o}K what will be the freezing point of 1000g of distilled water if 50g of the compound are dissolved in the water?

d. Regarding the first three elements (CHN) fill out the following table.

	C	H	N
No. of protons			
No. of Neutrons			
No. of e-			
Period on the periodic table			
Family on the periodic table			
Common oxidation number			
Outer Orbital formula			
Valence Number			
Principle (n) quantum number			

3. Name the element identified by the following formula.

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$

Ans__________________

4. Predict the molecular formula for a compound formed between Ta and S.

Ans__________________

5. Phosphoric acid (H_3PO_4) is one of the chemicals sometimes associated with shipboard water treatment. It is known that as a result of the solution process, the following occurs:

$$H_3PO_4\ (aq) \rightarrow 2H^+ + HPO_4^{2-}$$

Ans__________________

a. If 294g of the acid are dissolved in 2ℓ of H_2O what is the molarity of the soln?

Ans__________________

b. What is the normality of the soln described in part "a"?

Ans__________________

c. What is the mass of one equivalent weight of the acid?

Ans__________________

d. If the volume in part "a" above is diluted so that the final strength of the acid is 0.1N how much H_2O must be added to original 2ℓ volume?

Ans ______________________

e. How much 1M NaOH must be reacted with the 2ℓ volume in part "a" to neutralize the soln?

Ans ______________________

6. CO_2 from the atmosphere and the resperation of organisms dissolves dolomite (Ca CO_3) to increase the hardness of water.

$$H_2O + CO_2 + CaCO_3 \rightarrow Ca\ (HCO_3)_2$$

How many grams of CO_2 are required to produce 3 moles of $Ca(HCO_3)_2$?

Ans____________________

7. Carbonic acid (H_2CO_3), when present, in engineering plant water ionizes according to the following reaction to increase the H+ content which goes on to corrode steel boiler parts.

$$H_2CO_3 \rightarrow H^+ + HCO_3^-$$

If two moles of the acid are dissolved a liter of H_2O, (omitting a consideration of strong and weak acids, solubility product constants etc.) mass inconserved and the reaction goes to completion as shown?

a. How many moles of H_2 are produced?

Ans________________________

b. What is the pH of the soln?

Ans________________________

c. What is the pOH of the soln?

Ans________________________

8. In metal loss calculations, we know that:

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$$

If 32g H_2 gas are produced in a boiler tube each day now many grams of iron are converted to the corrosion product Fe_3O_4?

Ans__________________

9. Which of the following engineering plant water related equations are balanced?

a. $H_2O + CO_2 + MgO \rightarrow Mg(HCO_3)_2$

b. $2Na_2FeO_2 + 8H_2O \rightarrow 4NaOH + 4\,Fe_3O_4 + 16H_2$

c. $9N_2H_4 \rightarrow 12NH_3 + 3N_2$

d. $MgC\ell_2 + 2H_2O \quad Mg(OH)_2 + 2HC\ell$ Ans________________

10. Zinc in the presence of acid enters into the following corrosion reaction. If (NO_3) has a minus 1 charge:

$$Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NH_2(NO_3) + H_2O$$

a. Balance the equation
b. Identify the following

(1) oxidizing agent ____________

(2) reducing agent ____________

(3) element reduced ____________

(4) element oxidized ____________

11. Steam in a ships boiler at 600°F is found to contain the following elements/componds;

H_2O, O_2, CO_2 and H_2. The gases have the following partial pressures respectively 4 atm, 1520 TORR, 760 mm hg and 14.7 psi.

a. What is the pressure exerted by the steam on the walls of boiler.

Ans________________

b. If (R = 0.08205) there are 180g H_2O in the steam, what is the volume of H_2O present?

Ans________________

c. If conditions changed to STP and there were 2 moles of O_2 present, how much volume would you expect the O_2 to occupy?

Ans ____________________

12. The following scale forming reaction occurs in marine engineering plants.

$$Ca(HCO_3)_2 \xrightarrow{\Delta} \underline{CaCO_3} + CO_2 + H_2O$$

If a saturated soln at 100°C contains .0019g/mℓ $CaCO_3$

a. When the concentration of $CaCO_3$, exceeds the saturation value, that happens?

b. How many grams of $CaCO_3$ are dissolved in 10ℓ of plant water?

Ans ____________________

c. If the saturation value in g/mℓ decreases with decreasing temperature, is the soln exothermic or endothermic?

Ans ____________________

d. If this solution, saturated at 100°, decreases in temperature to 25°C where its saturation value is 0.0015 g/mℓ, how much $CaCO_3$ will precipitate per liter of plant water?

Ans ____________________